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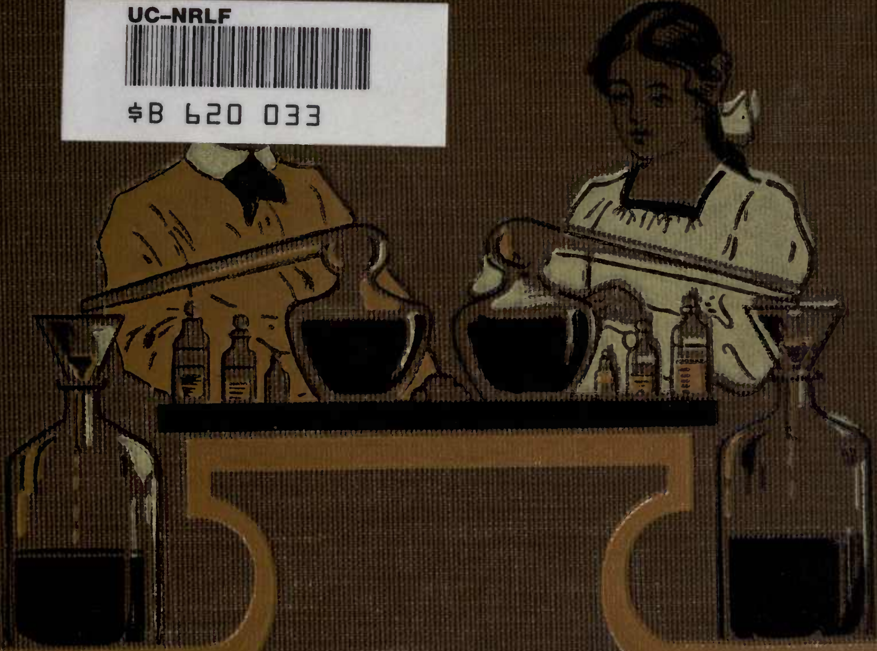

CHEMISTRY FOR YOUNG PEOPLE

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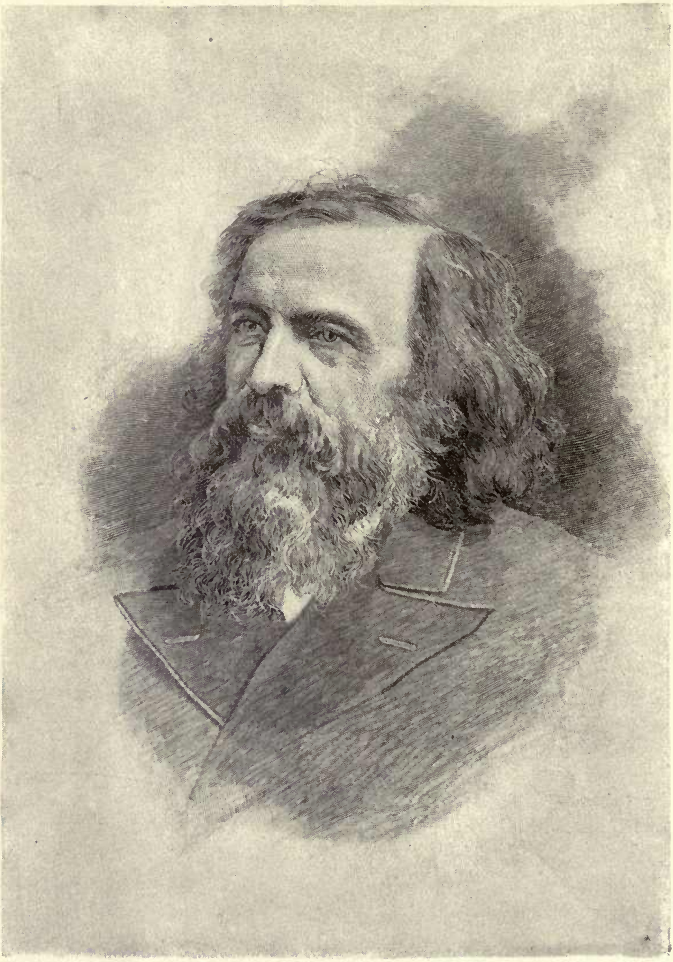

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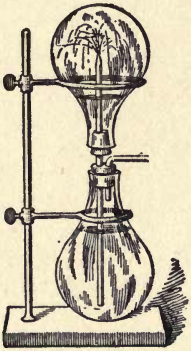
DIMITRI IVANOVITCH MENDELÉEFF
B. Siberia, 1834.

CHEMISTRY
FOR
YOUNG PEOPLE

By TUDOR JENKS

AUTHOR OF "ELECTRICITY FOR YOUNG PEOPLE,"
"PHOTOGRAPHY FOR YOUNG PEOPLE," ETC.

WITH SIXTEEN ILLUSTRATIONS AND
TWENTY-SEVEN DIAGRAMS



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INTRODUCTION

In purpose this book, like the two preceding volumes of the series, which dealt respectively with electricity and photography, makes no attempt to fill the place of a text-book, either in its contents or its purposes. It is planned to give the young reader such a general account of the modern science of chemistry as will prepare him to know better what place the science holds in modern life, and how it is related to other sciences. For this reason it is much more general in its nature than a school-book can be. It contains no complete account of any one of the elements, no directions for performing experiments or for laboratory work. It does, however, give an account, first, of the commoner elements and such a discussion of them as leads to a knowledge of modern ideas of what matter is, and of its laws—meaning, of course, by "laws" those general ways of action which experiments have made clear. When the reader has thus gained some acquaintance with the facts upon which modern chemistry is based, he is told of the broad principles governing all chemical actions, and learns something of the great philosophers and chemists from whom we have learned to make chemistry a science.

After this general acquaintance is secured, the commoner elements are taken up, and the reader is told those more important facts about each which should be known, even to one who desires no expert knowledge of their qualities. After obtaining such general knowledge of these elements, their qualities and their uses, the reader is led to enter a little more deeply into the greater principles that enable us to understand their action one upon another, and to become acquainted with the Periodical Tables wherein the relations of all elements to one another is shown to depend upon their atomic weight. These relations include the subjects of valence, of atomic heat, and, generally, of chemical action as classified by modern chemists.

To the enormous subject of organic chemistry, or the chemistry of carbon compounds, but one chapter is given; but in this are set forth such general laws and rules as enable the reader to comprehend the complexity of the subject and the methods by which chemists are able to make new compounds and to predict their qualities.

After this general review of the subject, there follows a discussion of the history of chemistry and of its relation to human progress, which contains, though in brief form, the important facts enabling young readers to know the importance of the part played by modern scientific chemistry in practical life. Then follows a set of tables which give in a complete way the essential facts in relation to all elements, thus completing those tables which in the earlier part of the book are confined to the more familiar elements.

It may be stated, again, that the object of the volume is to give the young reader such an interest in chemistry and such a knowledge of it as will show its relation to other branches of human knowledge, and will substitute in his mind definite and clear ideas of its main principles in simple form, in the place of vague and useless notions. It is believed that the reader will find in these pages all of chemistry that can be readily grasped and remembered by any except those special students who will necessarily go to technical books for practical information.

Even the skilled chemist cannot perform experiments without constant reference to text-books, and in all sciences to-day the student is compelled almost at once to acquire and to use a larger or smaller reference library. While technical books abound for the general reader who wishes merely an intelligent acquaintance with the main principles of any science, he is often compelled to acquire this knowledge by dint of combining extracts from perhaps a dozen volumes. To such a reader this volume should prove valuable.

TUDOR JENKS.

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Chemistry for Young People

CHAPTER I

FROM ALCHEMY TO CHEMISTRY

THERE are certain of the sciences with which we all know we are concerned daily. Some contribute so directly to our amusement, or are so universally taught, that it is impossible to grow up in entire ignorance of their principles. Even if we do not recognize that we are dealing with scientific facts, yet we learn the facts as children, and when we come to study the science to which they belong, find that many of its facts are already known to us.

One of these, for example, is the science of physics. Though the name may sound strange to a young student, yet he very soon finds that he is acquainted with many of the things taught him. He has learned that bodies have weight, are elastic, have cohesion or adhesion, may be measured and compared. In botany and in zoology every child has some little training. He learns to distinguish one animal and one flower from another, and when he is taught the science that relates to them, he finds he has learned already a great number of facts to which he can apply the principles he acquires.

With chemistry the case is different. It is one of the latest of the sciences. Although it is true that chemistry enters into every-day life, the changes with which it is concerned are not such as can be easily made clear. In order to comprehend them at all, it is necessary that there should be a wide acquaintance with a number of facts that are not readily understood, and indeed are not understood at all until certain principles have been learned. The evolutionary philosophers have taught us that children go through, in a way, the same stages that mankind went through in its general history; and as mankind was for ages ignorant of anything that can truly be called the science of chemistry, so boys and girls do not find it necessary to understand the facts of chemistry in passing through the stages of their life that correspond to the early stages of the life of mankind. The life of children, until they are well grown, does not require them to know more than is known to the savage races. It is at a much later period that they begin to put questions to answer which requires them to learn such general principles as those with which chemistry deals.

For many centuries mankind was satisfied with the general idea that all forms of matter were made up of different things put together, but it was only late in the history of mankind that the observation of the changes in substances led to the asking of questions that could be answered only by careful experiment, and it was still later that it became possible to construct the apparatus that made the experiments exact. It is only within about three hundred years, that is, since the sixteenth century, that there can be said to be any science of chemistry. Up to that time, there existed, even among the wise men, only certain general notions that had added little to men's knowledge since the days of the early Egyptians.

It is difficult to give an account of these early beliefs, especially because where there was so little exact knowledge, and where each wise man, magician, or priest—and these three capacities were often united in one individual—imagined for himself how substances existed, were made up, or acted upon one another, there could be little agreement, and their notions cannot be reduced into simple statements.

There was a general opinion that substances were to be roughly classed under four or five headings, that is, were made up of four or five "elements." These were said to be earth, air, fire, water, and ether. Air meant to the ancients every sort of gas or gaseous substance; fire was any sort of flame or heat, and as life seems always accompanied by heat, it came to be looked upon as the principle of life; water included all liquids, somewhat as air included all gases; earth was a general name given to all mineral substances, and, indeed, practically to all solids. Ether is more difficult to define. It seemed to include something that was more spiritual than air, and was used to explain what could not otherwise be accounted for. As the unknown region became smaller, there was less and less reference to the ether, and in later times the element, air, was used much as in earlier times ether had been. Even to Shakespeare's time we shall see that this general conception of the four elements prevailed among all save the most learned.

It was impossible that even the rude observations of early times should not show that now and again one of these elements seemed to be passing into another; for a solid when exposed to fire would first change to a liquid, and then would evaporate entirely; or, instead of evaporating, would give rise to a gas that caught fire and seemed to change to flame. Thus the very crudest experiments would show that the element "earth," or a solid, would change into the element "water," that is, a liquid, this into a gas or "air," which in turn would go back to flame.

From such observations it was certain that experimenters should be led to the conclusion that it was possible for one of their elements to take on the appearance, or to assume the qualities, of another. At a time when the imagination was active and was not controlled by facts (simply because facts were so little known), who can blame the old philosophers for their belief that the elements were capable of changing one into another, and thereby becoming convinced that it was possible to convert one into another, almost at will, if only the right method could be found?

From such conclusions came the beliefs of the alchemists, and in them we reach the beginnings of the science of chemistry.

In this name, chemistry, we have still the old Arabic word that appears also in alchemy. *Al* is, of course, the word for "the," as we see it in the name, Al-Koran, or The Koran; so alchemy simply means "the chemy." The origin of the latter part of the word is more or less obscure. So far as it has been traced, it seems to signify the extraction of the juice of plants for medical purposes, and thence to have come to mean the making of drugs, and processes akin to that. It therefore was early applied to the mingling and treating of natural substances to produce changes in them. In the course of time, those who were skilled in these arts often gave their efforts to the search for the means of making, or extracting, the precious metals, silver and gold. Believing that it was possible to change one substance into another, and not knowing enough to limit the possibilities of extracting one substance from another, the search for an easy means to make gold and silver was pursued by most of those who became skilled in laboratory work.

Knowing how futile were their efforts, we are likely to think of the old alchemists as engaged in a fool's quest. But before the rise of modern science, there was no reason why the ancients should consider the making of gold and silver out of other cheaper substances as a dream. Indeed, a few historians go so far as to admit that some stories recounting the success of alchemists seem to be supported by excellent testimony. But, at all events, the alchemists who searched for means to make gold and silver, and at the same time for means to prolong human life, made experiments and observations, learned methods and principles, that have been invaluable in founding modern chemistry. The history of these old alchemists is most fascinating, for, besides the actual practical facts of their experiments with all sorts of substances, and their studies into the laws of nature, there is a wealth of material relating to their dreams and fancies connecting the practice of alchemy with other arts and sciences, especially with astrology and astronomy. Many of the beliefs they held in regard to the connection between the heavenly bodies and earthly substances, have left traces in chemical terms and in words of daily use.

Each of the metals was supposed to be mystically related to one of the planets. Gold, as the most beautiful and least corruptible of metals, was assigned to the sun, the greatest object in the heavens; silver, for many reasons, was linked with the moon; mercury named alike the planet and the metal; the red and warlike planet Mars had its counterpart in the red-rusting weapon-metal, iron; and the qualities and virtues assigned by astrologers to the planets were believed to have their counterparts in the properties of these metals. The whole science, or art, of the alchemists has been well called "the sickly but imaginative infancy through which modern chemistry had to pass before it attained its majority, or became a positive science."

The ancient alchemists could go only a certain distance in their discoveries. They were patient, laborious, inventive, and were excellent observers. They were bold and imaginative in forming theories about their work, but they had not the power of bringing their theories to the test of proof, for they lacked the necessary apparatus to make their experiments exact and conclusive. This was one reason why chemistry came late as a grown-up member into the family of sciences. Astronomy, for example, took on the form of a science much earlier, since its facts required little else than observation and the applying of reasoning to what was seen. Long after astronomy had begun to be truly scientific, chemistry was still undeveloped, and was hampered by a mass of superstitious or magical theories that had not yet been cleared away.

In the sixteenth century there came about a change in the objects pursued by the alchemists. They had hitherto sought the "philosopher's stone" —a magic substance that would convert baser metals into gold; the "universal solvent," that would dissolve all substances; or the "elixir of life"—a drink that would prolong life. But, abandoning these wild quests, they now began to believe that their arts should be devoted to finding out how to prepare medicines.

The general terms used for the four elements about this time began to be replaced by more definite words. The chemists, or alchemists, of that later day replaced the old words, earth, air, fire, water, by what they called the "principles," salt, silver, and mercury. These were believed to be modifications of water, which was considered the true principle of all things, and air was considered a separate, somewhat unearthly element. About this period also grew up the recognition that the liquids had two general characters, being either acid or alkaline. It was believed that the health of the body consisted in a true balance of these opposite qualities.

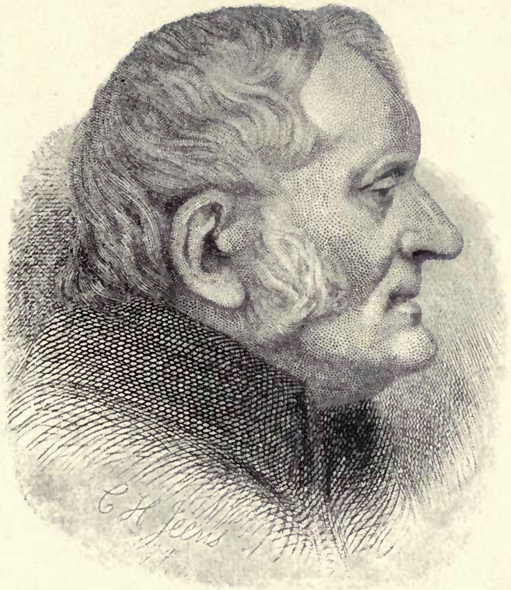
Robert Boyle, an Englishman living between 1627 and 1691, attacked these general notions, pointed out the influence of heat in forming new bodies, and suggested that all things may consist of one universal matter, portions of which differ from one another in certain qualities. This universal matter he believed to be capable of forming new bodies by a combination of its smaller particles or portions. But, acute as Boyle proved to be in pointing out the mistakes of others, it was impossible that he could then come much nearer to the truth than they had done.

Following Boyle came a number of experimenters who put forth various theories to account for what they had noticed in their laboratory work—of these we give a general account in a later chapter. It was not until the beginning of the nineteenth century that chemists were led by the researches of Dalton to what we now believe to be the right theory of the make-up of matter, the theory which put chemical work upon a true scientific basis and thereby led to the rise of modern chemistry. That modern chemistry possesses *all* the truth, is not to be asserted; but that it is practically true has been proved by the control which chemists have acquired over matter. They may not know all the facts in regard to how substances are formed, but at least they know enough to separate substances into their elements, and even how to make these elements join to make new and before unknown substances. Discoveries in the future may make this knowledge more accurate in detail, but it can never be entirely set aside. In other words, modern chemistry is true so far as it goes, and proves itself by its experiments. The old chemistry contained some truth, but there was little truth in the beliefs held by the older chemists to explain what they had observed. With Dalton, therefore, modern chemistry begins.

John Dalton, born September 5, 1766, in England, was a Quaker. He taught school, lectured on natural philosophy, and became interested in taking observations of the weather. For this purpose he made an instrument for finding out how much water there was in different falls of rain. But Dalton, not satisfied with measuring how much rain fell, took to asking the reasons why. He examined how much water was carried in the air under different conditions. Then he was led to inquire how the vapor of water could possibly exist in the air without becoming mixed in with the gases that make up the air. This led him to reason upon the question of the mixture of gases one with another, which led him to wondering how gases were made up; for, of course, each particle of any gas still remained *itself,* even when two or more gases became mixed together.

In Dalton's time it was already known that water consisted of the joining of two gases, hydrogen and oxygen. So the studious Quaker youth, when he began his keen reasonings, found himself obliged to account for the fact that the vapor of water could be mixed generally throughout the air, could then be deposited from the air, as in the falling of dew or rain; could evaporate, or be taken up by the air again, all without losing its identity as water. Yet it was possible by chemistry to separate this substance, water, into *two* gases, neither of which was like that from which it was separated, the liquid water.

When he had thought out this problem, Dalton



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JOHN DALTON
B. England, 1766. D. 1844.

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saw that there must be two *kinds* of fine division of matter. In one kind, although a substance was divided finely (for instance as into the particles of steam, which is the vapor of water), yet there was no loss of *identity*, each little, or littlest, portion was as truly the substance as the largest mass of it that could be got together. But there was another kind of division; for these finest portions could, if the proper means were used, be separated into portions which had entirely different properties. Then these smaller portions, no matter what was done to them, could not be further separated.

By means of this reasoning, verified by repeated experiment, Dalton came at length to believe that he had hit upon the secret of the joining of various substances into new compounds, and that this discovery made it possible to treat chemistry as exact knowledge.

We do not mean yet to explain how Dalton accounted for the making of substances by chemical laws; that we shall have to discuss in the book itself, and it is enough here to understand that with Dalton's discovery, made early in the nineteenth century, we may consider chemistry as first taking the right road, the road that has brought it to the control of fields of industry, of research, and discovery, greater, perhaps, than those belonging to any other science. And now that chemistry is truly a science, it can be brought into connection and relation with other sciences.

CHAPTER II

THE AIR. OXYGEN, OZONE. BURNING

WE ourselves cannot do better than to begin our knowledge of chemistry with the same substance that led John Dalton, the Quaker, to his discovery of the laws that underlie the science. As he began with thinking of the air and the raindrops, or water, so shall we. And though for a while we shall postpone the subject to study something of the gases that compose the air that surrounds us, yet it will be only that we may better understand how water is composed of two gases; for these two substances, water and air, properly studied, will tell us a great many of the main truths of chemistry, and something of the laws that have to do with all chemical combinations and with the elements that make them up, for water and air are each well known to be composed of two gases.

In the case of air, which we may well take first in order, these gases are simply *mixed,* not united into a new substance. They are the ones known to chemists as oxygen and nitrogen, the names meaning respectively acid-maker and nitre-maker. Exactly *why* animals and plants should have been created so as to need a constant supply of this mixture of gases, lies not only beyond chemistry, but beyond all the science we have at present; for like many another " why," this has to do with what is known as " metaphysics." It may be that creatures could exist, and possibly in other worlds do exist, whose life is supported by other gases, but so far as we can know at present, it is only by the use of this mixture of the two gases, nitrogen and oxygen, that every living creature carries on the growth and the changes that make up life. Each of the gases has its part to play. Oxygen may be considered the most active agent in nearly every change that takes place, but oxygen alone is not fitted to the uses of life. It has to be mixed with something that slows down its action, or life would soon cease. When we make an air that is slightly richer in oxygen, or go to such places as already have an atmosphere in which the oxygen is slightly more plentiful (as in mountain regions, or among certain plants that give it off freely) we find that the processes of life are hastened. Jules Verne has illustrated this in a fanciful story called "Doctor Ox"; and the same device of imagining people to live in an air very richly supplied with oxygen has been used upon the stage.

As has been said, by excess of oxygen life is hastened. Many of the changes that make up living take place more rapidly, and there is in the case of mankind an increase of vigor, a brightening of the spirits, with a quickening of the pulse, and a more rapid circulation of the blood. In the normal air something of the same results may be reached if rapid breathing brings more oxygen to the lungs in a given time.

In the case of birds, for example, the breathing is more rapid, so that the blood, by means of the lungs, receives more oxygen than in the case of mankind. Consequently their circulation is quicker, and many of their life-processes are more rapid and intense than in our own case. On the other hand, among the slow-breathing creatures, such as the reptiles, the precise opposite is true. Certain animals live at two rates of speed. The animals that hibernate, or go into a sort of retirement during the winter—bats, bears, and woodchucks, all being examples—there is a slowing down of life during their period of repose. They breathe very slowly, using but little oxygen, and their lives go on very much as a banked fire is kept up when the drafts are shut off.

In fact, that is precisely what takes place, for the fire itself is nothing but an example of the effect of oxygen upon matter. If we open the drafts of a fire and allow the lighter, heated air to be drawn through the burning material, the oxygen in the air makes the burning rapid and the fuel is soon consumed. When the drafts are shut off, and air is supplied in limited quantities, since there is little oxygen, the fire burns slowly, and if no oxygen is supplied, will cease to burn at all. We shall learn, when we come to study oxygen more closely, that what we call burning is known in chemistry under a different name. To burn is usually simply to *oxidize,* that is, to supply the gas oxygen to any substance that is capable of combining with it to make new substances. When combinations take place, we call the substance combustible; the chemists call it oxidizable. We shall, in fact, soon learn to consider every process into which oxygen enters as a "burning." This will include our own breathing, which is nothing more than a slow burning of the substances in the blood. A few combinations of other elements will also "burn," without oxygen, but only a few.

While it is true that pure air should contain mainly only two gases, oxygen and Nitrogen ~~hydrogen~~, as a matter of fact, the air to which we are accustomed is anything but pure. It always holds in it much moisture, various mixtures of other gases, certain very rare elements in small quantities, and also many forms of impurities. The atmosphere about us is really nothing but a great airy ocean, consisting, it is true, mainly of oxygen and nitrogen, but containing also, as the watery ocean does, an almost innumerable variety of gases and substances which are constantly entering into it in greater or less proportion, and then being removed, increased or diminished, according to circumstances.

One of the most important of these is "watery vapor," or water in a finely-divided state mixed throughout the air and held in suspension by it. When we for any reason wish to secure air that is perfectly dry and perfectly pure, it is necessary to filter it, just as we filter liquids, through fine screens or through cotton, and then to pass it through tubes containing certain substances which have the property of taking up the moisture from the air.

Those of you who have made photographs and used the platinum paper, have learned how difficult it is to keep the paper absolutely dry. It has to be stored in tin boxes, and in the box is kept a lump of a chemical substance (calcium carbide) to absorb the slight amount of vapor that from time to time must get into the box. But saying nothing at present about its impurities and mixtures, we will try to find out what we can in regard to the two gases themselves that mainly go to form a pure air. A full table of the air's composition will be found in the last chapter.

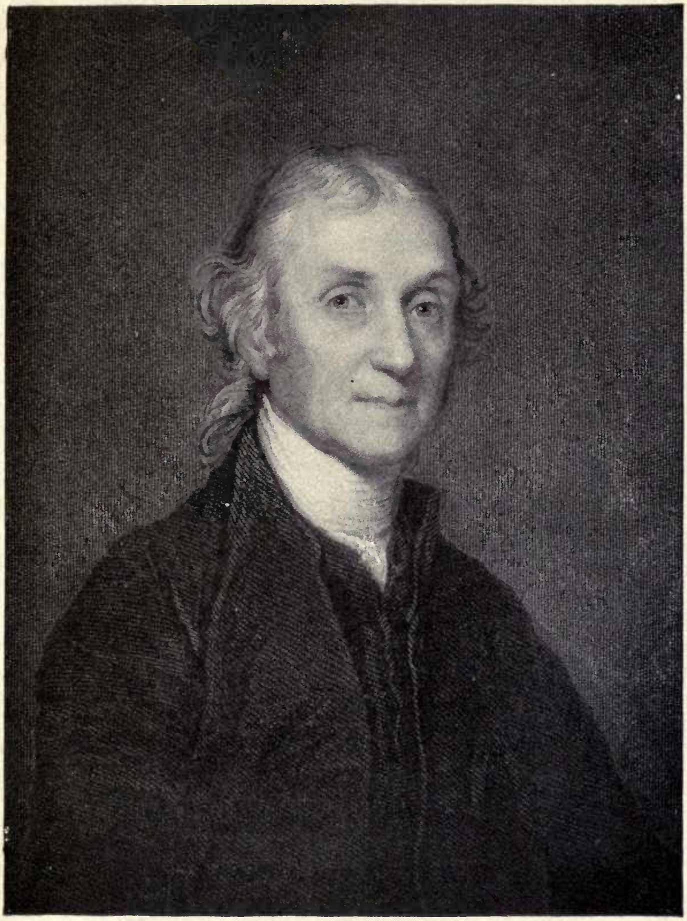
Let us mention in passing some recently discovered gases of which the air contains portions in minutest quantities—that is, minute in comparison with its own bulk, and yet enormous quantities if they could be separated from it and collected together. From Smith's "Inorganic Chemistry" we take a statement there quoted showing the elements that compose our atmosphere. In the order of their weights, we should have the following layers resting on the earth. . . . Five inches of water, thirteen feet of carbon dioxide, a mile of oxygen, four miles of nitrogen. Besides these there would be thin layers of the rare gases, argon, helium, neon, krypton, xenon. These gases, very recently discovered, seem to have no tendency to enter into any compounds, and so are of little importance in practical chemistry. The argon was first discovered, and then by treating it with liquid air, and heating it, the other gases are one by one caused to escape. It was only the power to produce these extremely low temperatures that led to the finding of these gases.

OXYGEN

The element, oxygen, is probably the best with which to begin the study of chemistry. It was discovered at an early date, it is more widely distributed and more abundant than any other element, and, especially, is found in the two great human needs, air and water. Of the air it forms about one-fifth in volume, and in water it is still more abundant, forming eight-ninths by weight. Almost half of the earth-crust itself is composed of oxygen compounds, and both vegetable and animal matter of every kind contain oxygen in greater or less degree. Thus it is not an exaggeration to declare that it is found everywhere and makes part of all substances with which we ordinarily have to do.

The discovery of oxygen was made by the English philosopher, Dr. Priestley, in 1774, and by the great French chemist, Lavoisier, independently. The name is Greek, coming from ὀξύς and γεννάω, meaning literally, "sour-maker," or "acid-maker." This name was given to it when it was found to enter into many compounds that if dissolved in water produced acids. The word "acid" means "sharp," and was given to these fluids because of their sour taste, and also, perhaps, in reference to their activity in attacking or dissolving other substances.

An easy method of obtaining oxygen by itself is that employed by Priestley in his earliest experiments. When mercury, or quicksilver, is heated, not to the boiling point, in contact with air, its surface becomes covered with a red scale. This scale comes from the combining of the oxygen in the air with the heated mercury, and it is called oxide of mercury. By collecting it from the surface of the mercury, and then heating it once more to a higher temperature, the mercury and oxygen separate, the first as a metal, the oxygen as a gas. If this second part of the experiment is carried on



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JOSEPH PRIESTLEY
B. England, 1733. D. Pennsylvania, 1804.

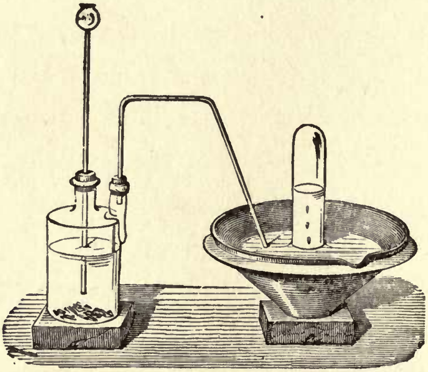
in a glass retort, or glass vessel, the oxygen rises to the upper part and may be drawn off free from the nitrogen and other gases that were mixed with it in the air. The metallic mercury forms in globules, or a film, on the cooler part of the glass retort.

It is just as well to explain here how this gas is collected from the retort. The retort may be, for example, a vessel globular in shape, with a long narrow tube projecting from the top. To the end of this tube is attached another tube of glass that goes into a pan of water. A jar is filled with water and then turned upside down, with the mouth below the surface of the water in the pan. The oxygen issuing from the end of the glass tube rises into the jar, and as it gathers, displaces the water. When enough has been collected, a flat piece of glass is slid over the mouth of the jar, which can then be lifted from the pan and remain full of oxygen.

This arrangement also was due to Priestley, and is called the "pneumatic trough."

A similar experiment may be made with other metals than mercury. With some the oxygen has little affinity, and from these it will readily part when heat is applied. These, of course, are the best for experiments.

Another experiment separates the oxygen in water from the hydrogen. This is done by conducting electricity into a jar of water, upturned as before over a pan, and providing another wire for the outgo of the current. As the current is conducted through the water from one pole to the other, it produces a separation of the two gases that are combined into a liquid. One of these gases gathers in bubbles at one pole, the other gas meanwhile gathering at the opposite. When enough gas is collected to form a fair-sized bubble, it will rise from the water. If, now, two small tubes have been



From Wurtz's "Elements of Modern Chemistry.

PRIESTLEY'S PNEUMATIC TROUGH

put over the ends of the wires, each will receive only one kind of bubble. One, therefore, will contain oxygen, the other hydrogen.

Neither of these methods is the one used for making oxygen commercially in large quantities, but the principle is the same in the commercial method. It consists of heating compounds of oxygen until the oxygen is given off as a gas. The compounds often used are potassium chlorate and manganese dioxide. Sometimes oxygen is got by forcing air through a tube containing a chemical that will absorb oxygen. From this compound the oxygen is afterwards recovered. A very recent method of obtaining oxygen is to condense air by cold and pressure into a liquid, when, if this liquid be allowed to evaporate, the nitrogen that was in the air escapes more rapidly than the oxygen, and the liquid left behind is almost pure oxygen.

When the gas has been obtained, it will be found to have these properties: it is colorless, odorless, tasteless; it is slightly heavier than atmospheric air. When oxygen is left in contact with water, the water takes up some of the oxygen, just as if the gas were dissolved in the water.

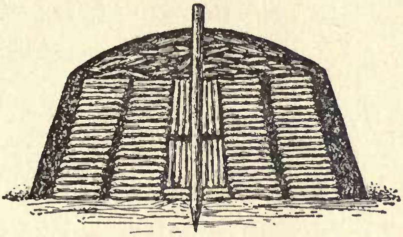
This last property is most important. Any water that is kept in agitation, as the ocean by its waves, or a river by its flowing, is constantly taking up oxygen from the air. The oxygen thus taken up supports the life of fishes, and also forms harmless compounds with all sorts of impurities in the water, thus making the water wholesome when drunk. Remember that although the water itself is by weight eight-ninths oxygen, yet *this* oxygen is so combined with the hydrogen in the water that it is of no use to the fishes in breathing. Neither is much of it available to combine with other substances, for it is not "free." It is the *free* oxygen added to the water that serves the fishes in breathing through their gills, and also acts as a purifier. The amount of oxygen thus dissolved by water is only about one part oxygen to thirty-three of water.

The most striking property of oxygen is the readiness with which it combines with almost all other elements. A few exceptions are fluorine, bromine, and the extremely rare gases already mentioned. This "combination" takes place in various ways. When exposed to the air metals show certain changes that result from the action of oxygen upon them. In the case of iron, this change is "rusting." The metal, iron, combines with oxygen to make an "oxide of iron," which is only the scientific name for rust. Lead, zinc, and copper, show a similar change to a less degree. A rust is formed on their surface, but since this rust is slight, and is noticed mainly by their becoming dull, we usually speak of these metals as being tarnished by air.

Phosphorus combines with oxygen very readily at ordinary temperatures. An instance of this property is the lighting of a sulphur match. The match consists of phosphorus and sulphur covered with a little paint. Scratching a match removes the paint, and slightly warms and compresses the chemicals, the oxygen of the air unites with the phosphorus, and the heat thereby produced sets fire to the sulphur, which in turn inflames the stick.

The heat and light caused by the action of oxygen is what gives the brightness and warmth gained in burning fuel. If a charred stick with the smallest spark upon it be thrust into a jar of oxygen, it burns fiercely. Even an iron wire, when heated to redness and thrust into oxygen, burns as if it were wood, throwing off sparks as it is consumed.

Except for the nitrogen in the air, the slightest spark would result in a most terrific and wide-spread burning of all sorts of substances within reach. The quick combining with oxygen, which we call burning, is not really different, except, in rapidity, from processes that we know by other names. The decay of wood and other substances, and the rusting of metals, are just as truly a burning, or oxidizing, as the actual flaming of a lighted match. After the process of combining with oxygen is finished, it is found that there is a change in the substances exposed to it, and the compounds produced by this oxidizing or taking up of oxygen are called, in chemistry, oxides. If these compounds are heavier than air, they are left as a result of combustion; if lighter than air, they rise, and are carried off, unless confined. Thus in an ordinary fire of coal and wood, the carbon (the element of which they largely consist and which is seen nearly pure in unconsumed charcoal) will, if plenty of oxygen is supplied to the fire, form a gas known as "carbon dioxide." The prefix di, is simply a Greek word for "two," or "double." The carbon dioxide, being lighter than air, is carried off. If a fire is burned without sufficient air to provide plenty of oxygen for its combustion, only a part of the carbon goes off in gas, and the rest remains in charcoal, or pure carbon. In the old folk-lore stories boys and girls will remember how often "charcoal burners" are spoken of. These people made their living by piling up wood into great heaps, covering it closely with clay, turf, or similar material, and then setting fire to it, but shutting off almost all the air. This was a method of separating from the carbon only such substances as would be easily consumed, and thus leaving the carbon, or charcoal, ready to be used again as fuel. The value of the charcoal as fuel in old times consisted in the fact that it made a clear, bright fire, burning slowly, and without much smoke—an excellent cooking-fire, before coal was known.



SECTION SHOWING WOOD ARRANGED FOR BURNING INTO CHARCOAL

It has already been said that the oxygen drawn into the body by the lungs is used somewhat in the same way as when it is admitted to a fire. It unites at first with the red corpuscles in the blood, and by these is carried throughout the whole body. It burns up the waste products of the whole system, thereby creating warmth and also getting rid of useless material. One product of this combining in the body is the same "carbon dioxide" that we find in fire smoke; and as that is allowed to escape through the chimney, so the carbon dioxide is brought back to the lungs and breathed out again into the air. Thus we inhale air rich in oxygen and exhale air deprived of much oxygen and rich in carbon dioxide. Rapid breathing increases the speed of these processes, but if carried beyond what is reasonable, produces a condition like that of fever. Indeed, fever seen in disease is only a condition wherein nature is trying to oxidize and get rid of an excess of waste material.

Although we have said that most cases of combustion mean oxidizing, it must not be thought that wherever heat and light are produced, oxygen is at work. The same effects may be produced by the action of other chemicals. Thus, for example, the gas chlorine combines so actively with hydrogen gas that there is a true "burning," which sometimes takes place so quickly as to be like an explosion. Sodium "burns" in chlorine, and phosphorus also; nor are these the only examples of the sort. Ordinarily, however, it is true that oxidizing is what we mean by burning.

OZONE

There is also, besides the ordinary form of oxygen, another form, in which the oxygen is, as it were, condensed. It is believed by chemists that this form comes about by a different arrangement of the particles of the gas by which they are able to be brought nearer together. This second form is known as ozone, a name derived from the fact that it has a striking odor, *ozo* being derived from the Greek verb "to smell." As you might guess from the name, while ordinary oxygen has no odor, there is a sharp and characteristic odor to the ozone. The idea of chemists is that in ozone three of the smallest possible portions of oxygen are grouped closely together, whereas in the ordinary form, only two of them are grouped. They are led to this conclusion for the reason that when oxygen is changed to ozone, it loses bulk and becomes only two-thirds as great in volume. In this way a hundred parts of ozone can be changed, or expanded, into a hundred and fifty parts of oxygen. This expansion is brought about by heating the ozone to a high temperature. To change it to ozone again, it is treated by what is known as the silent discharge of electricity—a discharge from a multitude of points. While oxygen has no color, a tube filled with ozone and looked at against a white light shows a slightly bluish tinge. All these matters will be better understood after the reader has gone over the fourth chapter of this book.

Both oxygen and ozone can be condensed to a liquid form by intense cold. In order to bring this about, ozone must be cooled to a temperature of more than a hundred degrees below zero Centigrade. Oxygen, to be reduced to a liquid, must be cooled to more than —118° Centigrade, and compressed under an enormous pressure. Ozone weighs more than oxygen, being quite a little heavier than air, while oxygen is only slightly heavier. There is another difference between the two forms of the gas. While oxygen is not easily changed from its usual condition, ozone is what is known as unstable, that is, it tends to change back into oxygen under ordinary conditions.

There are other reasons for knowing ozone to be condensed oxygen besides those given. While oxygen is an active gas, ozone, as we would expect, is even more active in oxidizing other substances.

This form of oxygen was first discovered by passing electricity through a tube of oxygen. Ozone is now known to be produced in the air during a thunderstorm. The odor suggests that of burning sulphur. As already said, its action is like that of oxygen, but more active in tarnishing metal, bleaching (by changing the composition of coloring matter) and corroding many softer substances. It has a use as a disinfectant. The air contains constantly a very slight trace of ozone, which one authority states is one part in 700,000. It is most abundant on the seashore and in the open country far from cities. A recent theory ascribes the formation of ozone in the air to the action of the "ultra-violet" rays of the sun upon the oxygen in the higher and exceedingly rare portions of the atmosphere. These rays are those most active in affecting the photographic plate.

We shall find, in considering other elements, that a number of them are like oxygen in appearing in more than one form. But that these forms are nothing besides the original element can always be shown by changing one form of the element back into another without trace of any other substance.

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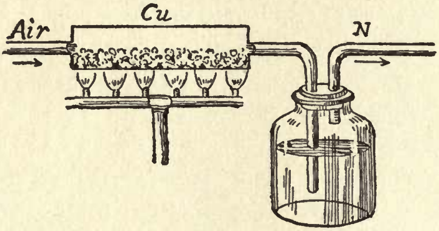
CHAPTER III

NITROGEN AND HYDROGEN

NITROGEN makes up four-fifths of the atmosphere in bulk. Its name was derived from the fact that it was found in saltpetre, another name for which is nitre. The gas, before it was known to be an important part of saltpetre, went by the name "azote," a word derived from the Greek and meaning that it was a gas which would not support life. This term is still used in some foreign countries, and also occurs in the adjective "azotized."

In the air the nitrogen and oxygen are not combined chemically, but are merely mixed together. As there is so much free oxygen in the air, nitrogen can be readily obtained by any means that will separate oxygen from it. One way of doing this is to pass a current of air through a tube containing red-hot copper filings or shavings, these being used because they give so much surface. When hot, the copper combines readily with the oxygen, making a copper oxide. Since the nitrogen does not so combine, it passes through the tube free from oxygen, and so cannot be collected in a suitable vessel.

When thus obtained, nitrogen is found to be a colorless, tasteless, odorless gas, as is oxygen, and slightly lighter than air. It will neither burn nor will it support combustion—that is, a burning substance thrown into nitrogen will, for lack of oxygen, soon be extinguished. If breathed by animals, nitrogen will cause death, not because it is a poison, but because the lungs being filled with it are deprived of oxygen.



Nitrogen, Prepared by Passing Air over Copper

Its principal use in nature, from our point of view, is its value in protecting us from the injury that would be done if the air were oxygen unmixed with this harmless gas.

The reasons why chemists assert that air, though made up of a mixture of nitrogen and oxygen, is yet not a compound of these gases, are these: If the two gases are mixed in the same proportions as those contained in air, we see none of the usual effects that come from chemical action. There is no heat, no change in bulk, and the mixture becomes like any ordinary air. Secondly, as will be seen later, when chemical compounds are made, the quantities of the various combining elements are always in fixed proportions, and the elements never combine except in fixed amounts. In the air, the proportion of nitrogen to oxygen is not that in which they thus combine. Third, in chemical compounds the proportions are always the same in the same compound, but in air the proportion of nitrogen to oxygen is not always the same. Fourth, and most convincing, air when shaken up with a quantity of water in a closed vessel, loses some of its bulk; then, if the water containing this dissolved air be put into another vessel, and be heated, it will expel this air, and the air will be found to be a mixture of nitrogen and oxygen in new proportions. Instead of containing *four* times as much nitrogen as oxygen, it will contain nearly twice as much oxygen as nitrogen. This would seem to show that the water has dissolved much of the oxygen and very little of the nitrogen. But if the air had been a chemical compound, the proportions of the two gases, when the air was recovered from the water, would not have changed.

If it be argued that the excess of oxygen may have come from the water, the answer is that in that case the water would have changed its composition, which is not the case.

Besides its forming so large a portion of the air, the gas occurs very commonly in the form of ammonia, a compound containing one atom of nitrogen to three of hydrogen. Another form in which nitrogen is found most useful is in nitric acid. Nitrogen may readily be prepared by covering burning phosphorus that is floated in a large vessel of water. The burning of the phosphorus is a combining of oxygen with it, and this oxygen being taken from the air in the jar, leaves nitrogen, more or less pure.

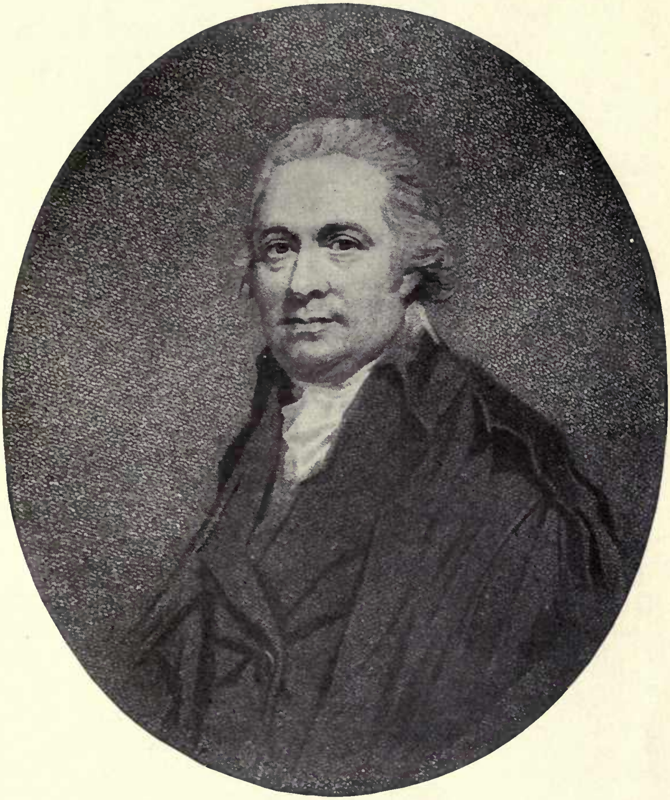
Nitrogen, as has been indicated in showing the proofs that air is a mixture, is not easily soluble in water. It requires a hundred parts of water to dissolve, or take up, about one and a half parts of nitrogen. Although the nitrogen in the air is not taken up by the system when the air is breathed, yet animal substances all contain nitrogen, a proof that they procure this gas indirectly. Most animal-foods contain compounds of nitrogen, and it is from this source that the nitrogen in the system comes. As animals procure their nitrogen compounds from food, so plants obtain nitrogen compounds from the soil, and if the soil be exhausted of its nitrogen, in order to make it productive it must be fertilized by the addition of nitrogen in some form. Not long ago it was discovered that certain plants, such as peas, beans, and clover, are able to take nitrogen directly from the air by the aid of bacteria, minute forms of life, found upon their roots. These bacteria provide nitrogen so freely that their roots serve to enrich the soil, and thus take the place of artificial fertilizer—that is, an exhausted soil can be renewed in its nitrogen by growing a crop of peas, beans or clover, upon it. Afterwards, plants that need the nitrogen compounds will find them in the soil so enriched.

Nitrogen was discovered in 1772 by Rutherford, a Scottish physician, who found that it would extinguish a candle, and suffocate small animals confined within it. Then the great French chemist, Lavoisier, showed what part it played in common air, and he it was who gave the name "azote."

In the branch of chemistry known as "organic," nitrogen plays a most important part, and besides its universal presence in animal flesh, gives rise to compounds of very marked properties—such as the albumens, the aniline dyes, the explosives, nitroglycerine and gun-cotton. In inorganic chemistry it gives us nitric acid, ammonia, and "ammonium"—which will be explained.

Having thus a slight acquaintance with the air and the gases of which it is a *mixture,* let us now take up some study of a chemical *compound*—the most familiar one and that Dalton studied.

As soon as we begin to examine the character of the substance, water, we find that we are dealing with something that is to be sharply distinguished in certain qualities from that mixture of gases known as air. Yet water, too, may likewise be separated into two gases. But precisely in the way it was proved that air is only a mixture, and not a compound (of oxygen and nitrogen), so it may be shown that water is not a mixture of two gases (hydrogen and oxygen), but is a true chemical compound of them.



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DANIEL RUTHERFORD
B. Edinburgh, 1749. D. 1819.

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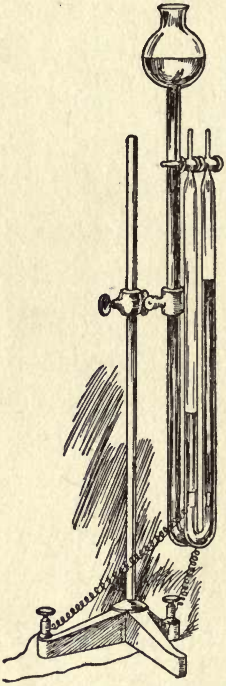
Yet water was considered, only a little over a hundred years ago, to be an element and not separable into anything simpler. The discovery of the true nature of water came about toward the end of the eighteenth century, and there were a number of steps before the nature of this compound was understood. Thus Priestley, the discoverer of oxygen, noticed that when a mixture of oxygen and hydrogen gases was exploded, there was a deposit of drops of water upon the cool sides of the tube. About a year later, another chemist, Cavendish, showed that when two parts of hydrogen and one part of oxygen were mixed and exploded, *nothing* was left in the vessel where the explosion took place, except water.

Two years later, Lavoisier, after going over the previous experiments of others, clearly recognized and explained just how the two gases, and they alone, when combined formed the compound, water. When this had been declared, other chemists proved and verified the statement. This they did by putting together carefully measured volumes of the gases, causing them to unite so as to form a measured quantity of water, and also by taking apart a quantity of water and securing from it the same proportions of hydrogen and of oxygen. Thus early in the nineteenth century it was finally proved that water was a chemical compound and always contained precisely two parts by volume of hydrogen, and one of oxygen. By weight there were eight units of oxygen to one of hydrogen.

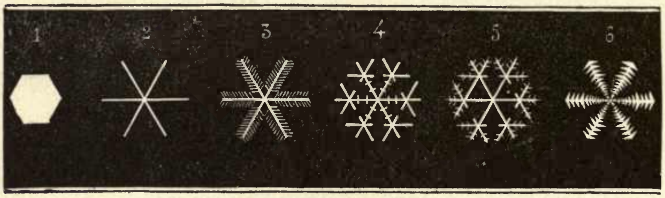
It is easy, nowadays, to verify these results. If the gases be mixed in a tube, the passing of an electric spark through the mixed gases causes them to combine, forming pure water. And, on the other hand, as has been noted, the electric current in passing through water will separate its gases and collect them at opposite poles.



HOFMANN APPARATUS FOR ELECTROLYSIS OF WATER.

The part which water plays in chemical actions is almost universal. It serves as a means by which the elements can reach one another and act upon one another. The combination of elements is for the most part brought about only in two ways, either by heat (or flame) or by moisture. When perfectly dry, most elements may be brought into contact without affecting one another in the slightest degree; but by "perfectly dry" is meant a dryness which can only be brought about by careful manipulation; and by "contact" is not meant a rubbing together with pressure—for this produces heat, and also may promote combination in other ways. Ordinarily there is in the atmosphere and in all substances more or less water or watery vapor.

We do not need to be told that the amount of this vapor, or water, varies very greatly, not only from time to time, but in different parts of the earth. We have every degree of moisture, from the parched and burning air of deserts, to the more than complete saturation of the air which results in the heaviest downpours of rain. It has been calculated that the average amount of water in a thousand parts of air may be put at about fourteen parts. Besides the enormous quantities thus held in the air, we know from our earliest studies in geography that three-fourths of the whole surface of the globe are covered with water. Even those substances which we ordinarily consider dry consist often for the most part of water. Every vegetable growth, the animals, even the rocks themselves, are filled with this liquid. The human body is about seven-tenths water, and most foods consist of very large percentages of water, from more than nine-tenths downward. We must not forget, when summing up all the amounts of water on the globe, the enormous ice-caps at the poles, the ice coverings of mountains, and their snowy summits.



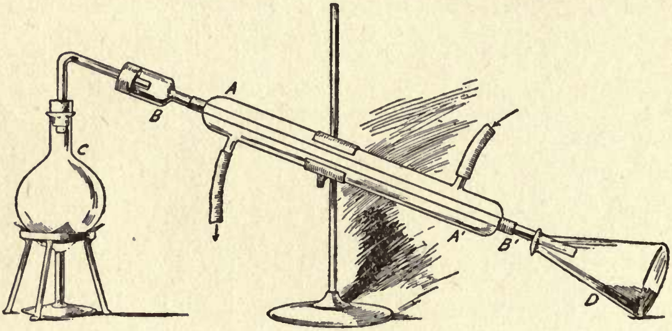
From Wurtz's "Elements of Modern Chemistry."
SNOW CRYSTALS

Thus the whole world may be looked upon as permeated and soaked in this combination of these gases, hydrogen and oxygen, and the liquid they form plays a multitude of useful parts in all the processes of nature. It will dissolve nearly all substances, and with them in solution flows freely from place to place, carrying them with it. It is constantly being taken up into the air, and as continually sent back again to the world beneath.

As it moves about on the surface of the earth it acts both chemically and physically. Chemically, it changes the substances with which it is in contact, taking away some elements and adding others. Physically, it changes the shapes and places of substances and their compounds, thus carving out the surface of the earth or in other places building it up, decomposes or forms rocks, and is, as it were, the life-blood of the world, causing in its surface and in its constitution changes similar to those that the blood brings about in animals.

Water also affects the substances with which it comes into contact by absorbing or giving out heat, thus changing their temperature. In its own transformation from solid to liquid, from liquid to vapor, and back again to liquid and solid, it not only affects the temperature of everything near it, but also takes up or parts with contained chemical elements, thus working other changes. By means of water, both plants and animals are enabled to take up and make use of the elements and compounds that are dissolved in it or carried along in its substance, and as will be seen later this "dissolving" is a process that is far from a simple one.

It is nearly impossible to conceive what our world would become without water, and by this it is not meant that we should imagine the world deprived of the two gases of which it is composed, but only of their combination as water in its ordinary liquid or vapor form. Most substances, when deprived of water, would crumble away into a powder. There could be neither animal nor vegetable life. A number of chemical changes upon which not only life, but the existence of numberless compounds depend, would altogether cease. Even where much of the water has disappeared from a



Condenser Arranged for the Distillation of Water

The condenser consists of an outer tube, AA, provided with an inlet and an outlet for a current of cold water, which surrounds the inner tube, BB. The vapor from the water boiling in the flask, C, condenses in the inner tube, owing to the decrease in temperature, and drops off from the lower end of this as the *distillate*, into the receiver, D, while the impurities remain behind in the flask.

heavenly body, as in the case of the moon, the result is to make it a dead and apparently useless body floating in space.

It is unnecessary to say to any thoughtful reader that the taking away of water from the substances at man's command, even if he could live without it, would stop nearly every species of human labor —beginning with all agriculture.

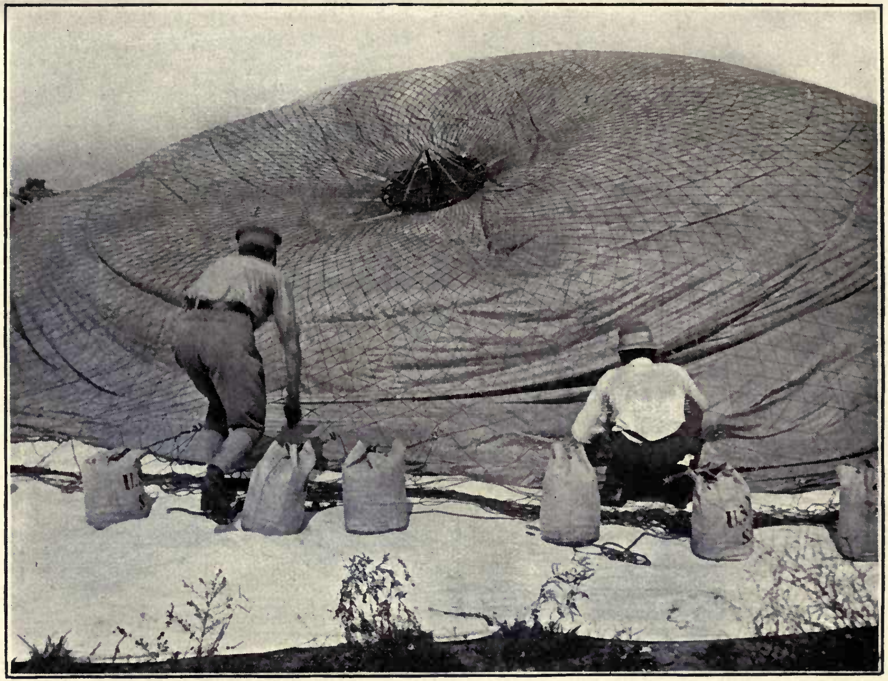
HYDROGEN

Of oxygen we have already told the few most important facts. But hydrogen, the other constituent of water, is even more interesting, as will be seen before even this elementary book is finished. There is indeed a theory that *all* things may be variations of this one element—hydrogen.

Of all substances, the element hydrogen is the lightest. Air is about fourteen and a half times as heavy as hydrogen; oxygen is sixteen times as heavy, and water eleven thousand times as heavy for the same volume. In our atmosphere, a mass of hydrogen floats rapidly upward, as a cork rises in water; therefore in pouring it from one vessel to another, it must be poured upward instead of downward.

A very ready method of showing the lightness of hydrogen as compared with air, is to use this gas for inflating a small balloon, or even for blowing soap-bubbles by means of a clay pipe thrust into a rubber tube connected with a reservoir of hydrogen. The balloon, or bubble, rises in the air with great rapidity. Hydrogen was formerly used for inflating balloons, but other gases, such as illuminating gas, are now found to be cheaper and better for the purpose.

When chemists decided, long ago, to arrange elements in tables with a statement of their relative weights, hydrogen was selected as the substance by which to measure the weight of all other elements. Just as in measuring lengths the metre or the foot is taken as the unit of measurement, so the weight of hydrogen is called the unit. The smallest possible amount of any element is called an "atom," and this word has long been used to describe a part of it so small that it could be separated further. Consequently, a single atom of hydrogen is taken to weigh "one unit of weight." Then by measuring the weight of any other element, or substance, or finding the number of times it will contain this weight, we may get a number that will express its



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INFLATING HYDROGEN BALLOON.

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weight as compared with hydrogen. In each case the amount considered is one atom. Since oxygen, for example, weighs sixteen times as much as hydrogen, when equal volumes of the two gases at the same temperature and pressure are compared, we may say that the atom weight, or atomic weight, of oxygen is sixteen, that of hydrogen being one.

A little thought will show the reader that this is as near as we can get to stating the weight of anything. We must always compare a substance with some standard and state that it is heavier or lighter, and so much heavier or so much lighter, than the chosen standard. There is no other way of telling what weight is. From this it follows that there is no answer to the question, What does the "1 unit" mean when taken as the weight of a hydrogen atom? We can only say that it means the weight of an atom of hydrogen—or if the weight of any other atom be taken as standard, then the weight of that atom.

If we take a given quantity of hydrogen, we can tell you what weight of any substance it is equal to, and we may state this in pounds, ounces, grains, or in the metric system as litres, decilitres, grammes, milligrammes, and so on. But in chemistry the weight of an atom of hydrogen was until recently always the standard. Latterly it has been proposed to take an atom of oxygen as the standard of weight, simply because this gives for other substances numbers easier to handle.

When we consider the make-up of various substances and find that the innumerable compounds in nature or in the laboratory can be simplified, analyzed, or taken apart, into a few elements, we are led to ask whether as we learn more about chemistry and become more skilful in reducing substances to their simplest form, we shall not at last be able to show that all things are different forms of a very few substances or elements, much fewer than the eighty now considered to be simple, uncompounded elements. We may in imagination go even further and ask whether we shall not come to a time when we shall know all substances to be only various forms of a single element. Many think, as already stated, that if this time should ever come, this single element will be hydrogen.

CHAPTER IV

PROPERTIES OF MATTER

THE first requisite in studying any substance is to get it into its simplest form. Matter is defined in books upon physics as being anything that can occupy space. Physics also teaches that all matter is divisible, or separable, into portions. But, for all that we can say, the defining of matter is impossible, except by telling what it does. Perhaps the easiest form of words in which we can put the thought is that which defines matter as being "anything that resists force." A little careful thought will show that we can know matter only where it resists some action of a force, as by being felt, seen, or otherwise known by the senses, by changing the action or direction of a force, as interrupting light or sound, and so on.

So far as we know matter, it is constantly divisible, and we have given names to these different degrees of division. If we take a substance and crush it to the finest powder, the properties of each particle remain unchanged. We may make these particles still finer by dissolving them in a liquid, or often we may melt them, or by heat convert them into a vapor or gas.

Take for an illustration the substance, water, which we have learned consists of hydrogen and oxygen. We may cause it to disappear by applying gentle heat, thus changing it to an invisible vapor; or, by a greater heat, we may convert it into steam, or by cold we may turn it into solid ice, and that ice may be melted again into water.

Yet we know by experiment that the tiniest particles of water may be brought back, unchanged, into a liquid form.

We know also, from chemical experiment, that the most minute of these divisions of water must contain two elements, hydrogen and oxygen.

For these different divisions of a substance chemists have made names whereby to follow substances in thought, even where they cannot be traced by the senses.

These names are as follows: To the smallest possible portions of a substance that still remain undivided into elements, and consequently unchanged in its properties, the name *molecule* has been given.

This word means, in Latin, "little body." For the still smaller portions that make up such a molecule, that is, in the case of water, for the tiniest portions of the two gases that go together to form the water molecule, the name *atom* has been made. Thus we say that the molecule of water contains atoms of hydrogen and oxygen. When we succeed in obtaining the smallest possible portion, or molecule, of a chemical element, or non-compound body, there is reason to believe that this too is made up of atoms, even though each of these atoms be the same element as the molecule.

We must remember, therefore, that all matter is considered to be divisible, first into molecules, and then into the atoms that go to make the molecule. This also implies that some molecules are made up of atoms that differ from it and from one another, while other molecules are made up of atoms that possess the same properties as the molecule.

The reason why chemists have come to think of matter in this way is because it explains to them and makes clear to our minds the different forms in which matter exists, and also explains how substances are made up chemically. The ordinary forms of matter known to all of us have already been mentioned in speaking of water, as solid, liquid, and vapor—ice, water, and steam. Nearly everything appears to us either as solid, liquid, or gas, and since we have seen that many common substances are very easily changed from one of these forms to another, we have come to believe that these result only from a change of condition.

Again, water is the most familiar illustration. We know that it is only a question of how much heat exists in water, whether we shall see it in the form of ice, or being converted to its melted form, as a liquid; and by still further heat we know it to be changed to a vapor.

But to explain these three states of matter it is believed that the strength of attraction holding the molecules together is what governs the state of the substance. When the molecules cling closely together and are not easily separated, the substance is a solid. When this force is much weakened so that the molecules, though somewhat held together, yet may be readily movable, the substance takes the liquid form. The force being still further weakened, the substance presents itself as a gas; and in this state either holds together with only the slightest union, or even tends to separate in all directions. The study of these conditions belongs not strictly to chemistry, but to physics. But a knowledge of them must be attained in order that we may follow the theories upon which chemists work.

When we come to dealing with the changes that affect the relations of the smaller portions, that is, of the atoms, we leave the science of physics and enter upon that of chemistry.

Chemistry has to do mainly with those changes in which the molecules are separated into atoms. Remember that we have said the molecule is the smallest possible portion of a substance that can retain its properties. If, now, this smallest portion be further broken up into atoms, we shall know it by a change in the properties of the substance treated.

When we shall have reached the state in which we deal with atoms, chemistry will show the laws by which these atoms go together again to make new compounds. By means of chemistry, then, we are able to do two things: First, to take apart the molecules of substances, and second, to put together the atoms which have been in these molecules so as to make molecules of new substances. The laws of the action of atoms in separating and combining, and the methods by which they may be taken apart and put together, and also the nature of their arrangement, and the force that causes them to remain together or to be separated—these are the field of chemistry.

There are in every science two great divisions. These are known as the "theory" and the "practice" (or, as they are sometimes called, the science and the art). The theory of any science is that part of it which forms the answer in any case to the question "Why?" The practice in the same way answers to the question "How?" If we find, for example, that by putting a fire under a vessel of water, the water gradually begins to boil, as we say, "boils away," we have learned something that relates to *practice.* We have learned how to change water into vapor. It is not necessary that we should know why the result is brought about, so long as we are satisfied with the result alone.

But as soon as we begin to wish to bring about any result in the best possible way, we must inquire why a certain course of action causes the result; and in the case of the water, we ask why heat should make water boil and then disappear. The answer to the question "How?" is usually a simple one. It can be found out by experiment. Once having found out, we may usually repeat the work as often as we choose. But the question "Why?" lies deeper, and sometimes cannot be answered at all. The answer to it is in all cases merely a guess—an attempt to explain more or less fully and satisfactorily. If we find that our explanation or theory makes it possible to foretell what will happen in new cases, then we may safely trust it and believe in it.

This whole matter of molecules and atoms is one of theory. None of our senses can enable us to know directly either molecules or atoms. We can only imagine that they exist, and then give reasons why their existence makes clear to us the action of elements or of compounds one upon the other.

It would be a long, dry, and difficult task to set out all the reasons why this belief in molecules and atoms—known as the atomic theory—has been accepted by nearly all men of science. It is not even yet complete or fully understood, but the belief in and knowledge of its laws have made it possible to do so many wonderful things that those who doubt it are few indeed.

This theory serves not only to explain many otherwise mysterious facts in chemistry, but also in physics, in medicine, in botany—in the whole circle of the sciences. It makes us see a reason for the rules and laws by which light, sound, electricity, and heat, act and may be controlled. It is desirable, therefore, to get into our minds a clear notion of matter, that is, of all substances, as explained by the atomic theory. Many great philosophers have told us good ways of picturing the nature of the theory to our minds. They ask us to suppose that a single drop of water could be enlarged until it were as big as our earth; then, it is believed, the molecule might be perhaps somewhere near the size of a tennis-ball. And these molecules themselves would be made up of still smaller atoms. This is not mere guesswork, but is based on careful figuring by Sir William Thomson (Lord Kelvin), who knew more of such matters than almost anybody.

Until very recently it was believed that the atom was indivisible and represented the very smallest portions of matter with which science had to deal. But especially after the discovery of the X-rays by Professor Rœntgen, careful study of their action convinced the scientific men that the ray was made up of a stream of moving particles of matter, much smaller even than the atoms. This came to be believed because of their going directly *through* so many kinds of matter. They would penetrate even sheets of metal. Then experiments were made to find out what these inconceivably minute particles weighed, and experiments led to the conclusion that they weighed but one-thousandth as much as a single atom of hydrogen.

It is, of course, not to be expected that in a book of this sort these complicated experiments can be fully explained. We will say only that the measuring and weighing of the velocity and the mass of these particles was done by causing them to move aside from their regular lines of motion by means of a magnet. Then it was calculated what their velocity and weight must be in order to be bent a certain amount from their paths by the known attraction of the magnet.

So, to-day, we have come to believe that from the atom itself can be separated something still smaller. Sir Oliver Lodge, another great English philosopher, has told us something of what these minutest forms of matter are thought to be. The name given to them is "electrons." If we imagine an atom enlarged to the same enormous degree as we have already imagined the enlarging of a molecule, then, Sir Oliver Lodge tells us, it would be found to consist of a few electrons—that is, few in proportion to the space they would then fill—flying about in orbits much as the planets move about the sun. Supposing that the atom had been enlarged until it were as big as a great house, then the size of each electron might be compared to that of a printed period, or the smallest dot that you could make with a sharp pencil-point.

If the atoms in a single drop of water are so tiny that enlarging a drop of water to the size of this earth might make each atom no larger than a tennis-ball; and then if we could imagine each of these tennis-balls magnified until it were as big as a church, in which case the electrons it contained would be no more than the tiniest dot—you may imagine for yourself the inconceivable smallness of these electrons, hundreds of which are found in every atom. The pencil dot you made to *represent* an electron must in reality contain millions of electrons, yet you must not think of these as packed closely together. They are, in proportion to their including space, as far apart as the planets in the solar system. They are in motion with unimaginable rapidity.

Exactly in what paths they move we cannot tell. They seem to be repelling one another and attempting to get as far apart as possible. It is believed that they are kept together, that is, kept from flying outside the limits of an atom, by electricity. It is believed, for reasons too long to give here, that all the electrons are negatively electrified. Now, like electricities repel, and so electrons repel one another. But it is believed that within each atom there is *something* positively electrified, which, therefore, attracts the negative electrons and keeps them revolving about it, just as the sun attracts the planets and keeps them from flying off into space. Whether the electric forces of which we are talking are connected with some portion of matter, or whether they consist of nothing except the motion of the *ether,* is not yet known.

To understand what is meant by motions of the ether, let us suppose a great body of water in which there are numberless and tiny whirlpools moving with enormous force. You know that the water coming from the hose of a fire-engine or of a mining hose travels so fast that it cannot be penetrated.

If you try to strike a stick through one of these streams, the stick will rebound as if it had hit metal. If you imagine the currents of the little whirlpools travelling with something like that rapidity, you will see that you could not push something into them. They would indeed resist as if they were solid. Now, it is believed that the electrons are in swiftest motion within each atom of a substance, and that they resist the attempt of anything to come between them, just as a stream of water resists the entrance of a stick. Sir Oliver Lodge, in speaking of the atom of hydrogen, the lightest of all, supposes that it may contain perhaps seven hundred electrons, half negative, half positive. If this number may be accepted, then an atom of oxygen, which is sixteen times heavier, might be thought of as containing eleven or twelve thousand electrons, and so on with the other elements, each containing as many electrons as its weight would indicate when compared with the weight of hydrogen.

He also warns us that though science has detected these negative electrons in certain rays, the *positive* “ electron ” is only imagined, having never been detected separate from some atom. This same theory carried out would lead us to believe in the possibility that all elements are made up of the same kind of electrons, differing only in their number, and possibly in their kind of motion.

Remember that all this is only a belief; but it is a belief that helps us to understand and explain by the laws of electricity alone many of the actions of elements on one another, chemically, that would otherwise be very complex and mysterious. Thus it is believed that all light comes from the setting free of negative electrons. They pass out from their atoms and shoot through space. When they are released from atoms with slower speed, they become evident to us as heat. When released in another way, they are known as electric action. By the same theory—the escape of electrons and their striking against other substances—is explained what is known as radio-activity, or the emission of different forms of rays, as from radium and other substances.

To show how recent these views are, it was only in 1903 that scientific men decided radio-activity was likely to consist in the flinging away of atoms positively electrified. These atoms have been weighed and seem to be somewhat near the weight of an atom of hydrogen, and they are accompanied by the negative electrons, only $\frac{1}{1000}$ as large.

One need not try to follow too closely the complicated thinking of the philosophers on these subjects. It is enough that we bear it in mind that all matter is believed to be made up, first, of *molecules* representing its own smallest portions; that these molecules are made up of *atoms* closely held together, and that the atoms themselves are made up of *electrons* moving in orbits with inconceivable rapidity round about some central portion. Such is the latest view of the make-up of all substances, and to it we must now and again refer in order to explain chemical laws.

In talking of the three forms of matter we have explained how, as heat is applied, a solid is changed to liquid, and the liquid to a vapor or gas. But what we have been explaining in regard to the make-up of molecules and atoms prepares us to understand what is meant by philosophers and chemists who nowadays talk of the " fourth state " of matter. In a gas the molecules tend to separate from one another. If this separation be carried still further, so that the molecules themselves, or even the atoms of a single element, are broken up into the electrons, then we shall have a new, or a fourth, state of matter, with laws of its own differing from those of the other three. Concerning this fourth state, we know as yet less than of the other three, and it is at this time being most busily studied in many laboratories.

It is especially interesting because to this state of matter belong the various new kinds of light rays whose action is so wonderful and was until the last few years so unaccountable. As a result of this whole method of thinking of matter under the atomic theory, we must regard the world, and the universe itself, as being made up of almost unimaginably tiny portions, each of which is eternally in motion of inconceivable rapidity. Thus it will be seen that the old idea of "dead matter" has been entirely abandoned. Everything is in motion, and science is little more than the study of these movements.

We may also, as a result of these theories, understand why it is that some scientific men believe that all the other elements may be nothing more than different forms of hydrogen; and for this reason. Hydrogen is the lightest of all matter. If we are trying to find that of which all other matter is made up—taking it for granted that there is one such universal element—then it is hard to escape the conclusion that this one element must be that of which the atom is the lightest of all things; for surely this atom cannot be made up of larger atoms if each one of them is heavier than itself.

To repeat the argument: Admitting all things have weight, if any one thing be *simple* and all the rest *compound*, every conceivable compound must be heavier than the one simple thing, since whatever is *added* to that simplest and lightest substance must make it heavier. If this reasoning be just, then it would seem hydrogen is that element which is the foundation of all the others.

In the same way it may be argued that since electrons are lighter than the hydrogen atom, they may be either the simplest form of matter, or something different from matter. Then at a certain point, as matter is being separated, it ceases to be matter when it is separated into electrons. Of course, to say that electrons are not matter is to consider them only as forms of motion in the ether.

We have already said that while the science of physics deals with molecules, or with substances whose properties do not change enough to change their identity, chemistry deals with atoms, and studies substances having regard to the atoms of which they are made up, and the changes in substances that are made by changing the atoms that compose them. It must be remembered, however, that even in chemistry we deal ordinarily with substances in molecules, since the changes brought about by separating and combining atoms are not possible to observe. The atoms do not remain in a free state. No sooner are they set free from one combination than they enter into another.

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CHAPTER V

THE ELEMENTS. THE LAWS OF COMBINATION

IN order that record may be kept of these changes of atoms, and also that they may be understood and followed, there is a system of notation and of naming. The chemical elements have each assigned to them a simple symbol. Ordinarily the symbol of an element is the first letter of its name, but where this would make confusion, other letters may be added. Some of the older elements are still called by their Latin names, and the initial is taken from the Latin word. Each of these symbols stands not only for the element, but for one atom of that element. If we wish to express more than one atom, a number is put before the symbol.

To take as illustrations elements of which we have already learned something, oxygen has the symbol O, which means one atom of oxygen. For nitrogen in the same way we have the symbol N, and for hydrogen the symbol H, the letter standing in each case for a single atom. If we wish to express more than a single atom, we put a number before it. Thus 4 O would stand for four atoms of oxygen. When the atoms are in *a compound,* the figures expressing their number is written *after* the symbol and may be made smaller and either higher up or lower down. Thus the well-known symbol for water is H_2O or H^2O, meaning two atoms of hydrogen combined with one atom of oxygen because water is so composed. Students of algebra will see that this sort of notation is like that in algebra, being no more than the use of co-efficients and exponents.

Having this way of writing the names of elements in compounds, it is easy to express any substance chemically. To do this we have only to put the chemical elements that enter into it side by side. There is no rule as to the order in which the elements are to be written, but certain orders have become usual. Thus OH_2 would likewise mean water, but it is not commonly so written.

Just as a symbol represents one atom, *so a formula represents one molecule.* To express more than one molecule, the number of them is written before the formula. Thus to write four molecules of water, we put it thus: $4H_2O$. Where no figure follows a symbol in a formula, the figure 1 is, of course, understood. Thus HCl would mean one atom of hydrogen combined with one atom of chlorine, and the whole formula means: one molecule of hydrochloric acid, as will be understood later.

Certain methods that are used in algebra may be extended to the writing of formulæ in chemistry. For example, if a compound acts as if it were an element, this is sometimes expressed by putting it into a parenthesis, or by separating the formula into groups by inserting a period. If a group of atoms is to be repeated several times, it may be enclosed in a parenthesis and a small figure put after it, just as if it were a single symbol. If the whole formula is to be multiplied, remember that the large figure precedes it all.

In order to express the fact that certain chemical actions take place, it is usual to write them in equations, which are read exactly as if they were arithmetical or algebraic—that is, from left to right. As the formula of any compound shows the number of atoms that are combined to make a molecule of that compound, and as no atom can ever be destroyed in making new compounds, a correct chemical equation *must always be equal on the two* *sides;* that is to say, if the number of atoms on one side are added together, they must equal the number of atoms on the other side. Likewise, if we add together the weights of atoms in any compound, we shall have the weight of its molecule, since the formula represents one molecule. In an equation, likewise, the weight of the molecules put together must equal the weight of the molecules produced.

In order, therefore, to use chemical equations, it is necessary for us to know the weight of each of the atoms. To the finding out of these weights for each of the elements has been devoted labor unimaginable, by countless chemists; and many of them are still somewhat uncertain. But for most of the elements, they are wonderfully accurate.

Three tables, giving important facts about the elements—first, a list of those that should be very familiar, then of those less well known, and then of rarer ones, are given on pages 74 and 75. Besides these there are many still rarer, which will be found in a complete table at the end of the book.

This writing of chemical equations and the expressing of chemical compounds by means of letters and numbers has been of the utmost value. It has led men to record and to remember the make-up of compounds, and, still more important, has led to the discovery of the laws by which compounds are made. It was thus ascertained that when a compound was separated into its elements, these elements were always found to exist in precisely the same proportion in the same compound. Thus in the case of common salt, chloride of sodium, as it is called chemically, no matter how much or how little salt was separated into sodium and chlorine, the *relation* of the amounts of these two to one another never changed. Sodium chloride has this formula, $NaCl$, and from what we have already said, you will understand that each molecule is made up, therefore, of one atom of sodium to one of chlorine. But the atomic weight of sodium is 22.88, chlorine is 35.18, and the weights of the sodium and chlorine in any amount of salt will always be to one another in the proportion that these numbers bear. Thus in 58 ounces of salt, there would be a little over 22 ounces of chlorine and a little over 35 ounces of sodium. If there were 58 tons, or 58 grains, the *proportions* of the two elements would not be changed, and if, on the contrary, we were to put together sodium and chlorine to make salt, the two elements would combine

TABLE OF COMMONER ELEMENTS

NAME	SYMBOL	VALENCE	STATE	OCCURRENCE	AT. WT. O-16	AT. WT. H-1	SPECIFIC GRAVITY	DATE AND DISCOVERER
Hydrogen	H	I	Gas	Gas, Water	1.008	1.	.00088	Cavendish, 1766
Nitrogen	N	V	"	Air	14.04	13.93	.00124	Rutherford, 1772
Oxygen	O	II, VI	"	Air, Water	16.	15.88	.0014	Priestley, 1774
Chlorine	Cl	I, VII	"	Salt	35.45	35.18	.00317	Davy, 1810
Aluminium	Al	III	Metal	Clay	27.1	26.19	1.71	Wöhler, 1828
Iron (ferrum)	Fe	II, IV	"	Ores	55.9	55.5	8.	prehistoric
Nickel	Ni	II, VIII	"	"	58.7	58.3	8.8	Crondstedt, 1751
Copper (cuprum)	Cu	I, II	"	"	63.6	63.1	8.9	prehistoric
Zinc	Z	II	"	"	65.4	65.9	7.12	Paracelsus, 1520
Silver (argentum)	Ag	I	"	"	107.93	107.12	10.5	prehistoric
Tin (stannum)	St	II, IV	"	"	119.	118.1	7.3	"
Platinum	Pt	II, IV	"	"	194.8	193.3	21.5	Wood, 1741
Gold (aurum)	Au	I, III	"	"	197.2	195.7	19.3	prehistoric
Lead (plumbum)	Pb	II, IV	"	"	206.9	205.35	11.	"
Mercury (hydrargyrum)	Hg	II	Liquid	"	200.	198.5	13.6	"
Carbon	C	IV	Solid	charcoal diamond graphite	12.	11.91	3.	"
Sodium (natrum)	Na	I	"	Salt	23.05	22.88	.97	Davy, 1807
Magnesium	Mg	II	"	Earths	24.36	24.18	1.75	Davy, 1808
Silicon	Si	IV	"	Quartz	28.4	28.2	2.5	Berzelius, 1823
Phosphorus	P	V	"	Bones	31.6	30.77	1.8	Brandt, 1738
Sulphur	S	II, VI	"	Volcanos	32.06	31.83	2.07	prehistoric
Potassium (kalium)	K	I	"	Ashes	39.15	38.86	.87	Davy, 1807
Calcium	Ca	II	"	Lime Chalk	40.1	39.8	1.58	Davy, Berzelius, 1808
Bromine	Br	I, VII	"	Seaweed	79.96	79.36	3.19	Balard, 1826
Iodine	I	I, VII	"	"	126.85	125.9	4.9	Courtois, 1811

TABLE OF ELEMENTS (FREQUENT)

NAME	Symbol	Valence	State	Occurs	At Wt O = 16	H=1	Sp G	Date and Discoverer
Fluorine	F	I, VII	Gas Non-Metal	Minerals	19.	18.9		Ampère, 1810
Arsenic	As	V	Metal	Ores	75.	74.4	5.76	Schröder, 1694
Antimony (Stibium)	Sb	V	Metal	Ores	120.2	119.3	6.71	Valentine, 1450
Bismuth	Bi	V	Metal	Minerals	208.5	206.9	9.82	Valentine, 1450
Cobalt	Co	II, VIII	Metal	Ores	59.	58.56	8.6	Brandt, 1783
Lithium	Li	I	Metal	Minerals	7.03	6.98	.59	Ariordson, 1817; Bunsen [1855
Manganese	Mn	II, VII	Metal	Minerals	55.	54.6	8.	Gahn, 1774
Strontium	Sr	II	Metal	Minerals	87.6	86.94	2.5	Davy, 1808
Barium	Ba	II	Metal	Minerals	137.4	136.4	.40	Davy, Berzelius, 1808
Boron	B	III	Non-Metal	Borax	11.9	10.9	2.	Davy, Gay-Lussac, 1808
Cadmium	Cd	II	Metal	Zinc Ores	112.4	111.6	8.65	Herman, Stromeyer, 1817

RARER ELEMENTS

Iridium	I	II, IV	Metal	with Osmium	193.0	191.5	18.7	Tennant, 1804
Chromium	Cr	II, VI	M & NM	Iron Ore	52.1	51.7	5.	Vauquelin, 1797
Osmium	Os	II, VII	Metal	with Platinum	191.0	189.6	10.	Tennant, 1803
Palladium	Pd	II, IV	Metal	with Platinum	106.5	105.7	11.4	Wollaston, 1803
Radium	Ra	?, II	Metal	Pitchblende	225.0	223.3		Curié, 1898

For the complete table of Elements see last chapter of this book.

with one another in these same proportions only. If there was more of one or the other, it would simply be left uncombined, for lack of a partner or set of partners.

This is what has become known as the Law of Definite Proportions. The law of definite proportions was found by Proust. It was seen that it did not in all cases explain the action of the elements, for at times the elements, though combining in definite proportions, combined in several proportions. By further observation, however, it was discovered by John Dalton that when an element combined in more than one proportion, the larger proportion was always a multiple of the smaller. Thus, supposing that the oxygen atom be represented by its combining weight, 16, when it combined in other proportions than 16 it was found that these were multiples of the number 16; that is, it combined in 16, in 32, in 48, or in 64, parts, as if 1, 2, 3, or 4 of its atoms combined with other elements as if they were but a single atom. When this law was recognized by Dalton, it became known as the Law of "*Multiple* Proportions." *

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* In the combining of gases, advantage is taken of the fact that they have been discovered to combine in "volumes in simple

This it was that laid the foundations of modern chemistry.

Upon these laws scientific study of all compounds were based, for with them combination became a matter of fixed laws, not of untraceable chance.

We shall now return to our study of the elements, taking up next that most important one, carbon, because of its abundance.

ratios to each other and to the volume of the product." That is, their relation is expressed by small whole numbers. Their volumes also are believed to contain, at the same heat and pressure, the same *number* of molecules. Thus they are compared with hydrogen, and weighed volume for volume. This gives the ratio of their weight to hydrogen—as will be explained later.

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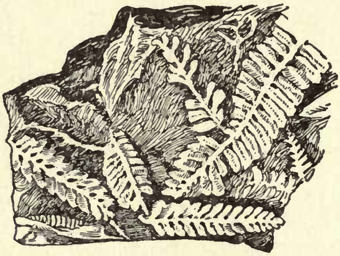
CHAPTER VI

COAL AND CARBON

COAL is so important a part of our civilization that, when only a few years ago, a strike on the part of the coal miners at that time put an end to the production of coal and caused it to become very scarce, it seemed as if there was no part of our daily life that did not depend upon it. Yet it is a substance that was almost unknown a few generations ago. It is true that in books written about that time we see occasional references to the use of coal in an open grate, to make a luxurious fire, but it was not looked upon as a necessity. Ordinary heating was done, where it was done at all, by burning wood. The work of the world was performed either by man or animal power, or by mills driven by wind or water. It was the increasing scarcity of wood and the increasing value of the steam-engine that suddenly gave to the deposits of coal an enormous value and set mankind to getting it out of the earth in great quantities.

Being so much used, it was also more deeply studied, but in spite of all research that has gone to determine its origin, and of all the experimental science bent upon it, coal even to-day remains more or less of a mystery. The earliest idea considered it a mineral, which, of course, in a sense it is, but not in the sense that granite or flint is a mineral. It is, rather, a mineral in the sense that chalk is so-called—that is, it is the remains of animal or vegetable life.

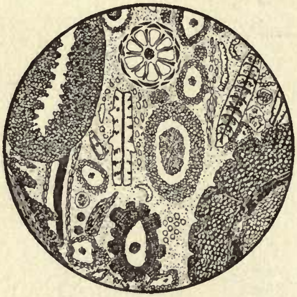


Fossil Found in a Coal Bed

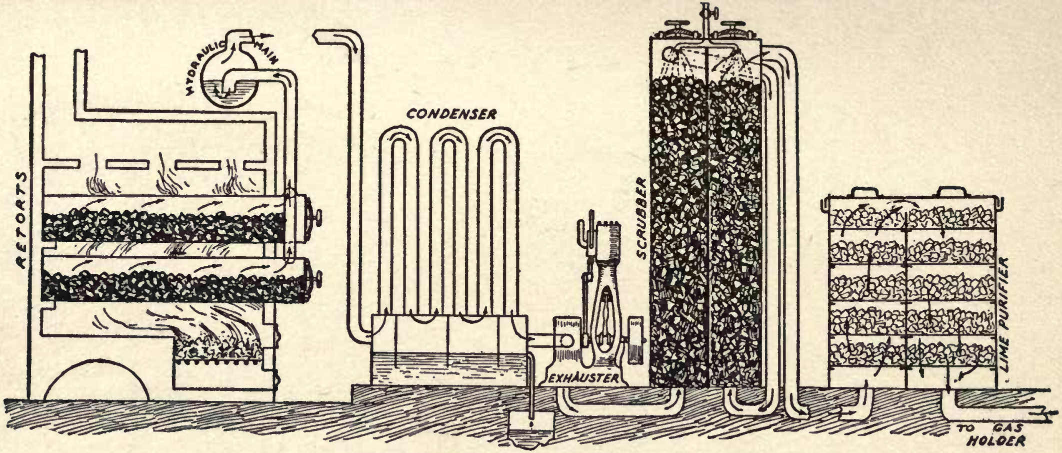
The next belief that man held about it saw in it the compressed vegetable substances of past ages—the wood and the leaves of old forests. This was a step nearer the truth, or what we now hold to be the truth. For, undoubtedly, coal consists of the remains of former vegetation, but instead of being mainly the product of an age wherein flourished great forests such as we now have, it is believed that coal is made up for the most part of the pollen dust that for thousands upon thousands of years, during the carboniferous age, was deposited from the gigantic ferns and similar plants that grew with amazing luxuriance in those hot, moist, and absolutely windless ages.

Whatever may have been its origin, to the chemist coal is a veritable storehouse of treasures. It consists, it is true, to a very large extent of carbon —the same substance that in other forms we know as the "black lead" of pencils (which nowadays is graphite, a nearly pure carbon) and in the diamond, which is simply pure carbon in crystal form. But, in addition to its greater proportion of carbon, coal has given rise to a vast number of other products. By being raised to a high heat in a closed vessel, the coal may be separated, or distilled, so as to give off fumes from which may be obtained various compounds and elements.

We have already said that coal is a vegetable product, and that it consists largely of carbon. And the same is true of all vegetable matter. Carbon is the foundation upon which all animal and vegetable tissues are built up. Owing to a strange peculiarity of the element, carbon, it has the ability to enter into an almost unending number of compounds. So complex are these and so puzzling, that in the science of chemistry the study of carbon compounds has been set apart as a separate branch. For many years, in the earlier days of the science, it was believed that a large part of these compounds could not be formed in the absence of animal life; that is, they must be the result of vital forces. It was therefore thought they were produced only in the tissues of living animals or vegetables. At a later date, however, chemists began to form these substances in the laboratory, and what had been called "organic chemistry" (that is, the chemistry of organic life) was studied under a new name—the "chemistry of carbon compounds."



Section of Coal as Seen Through a Microscope



APPARATUS FOR MANUFACTURING COAL GAS

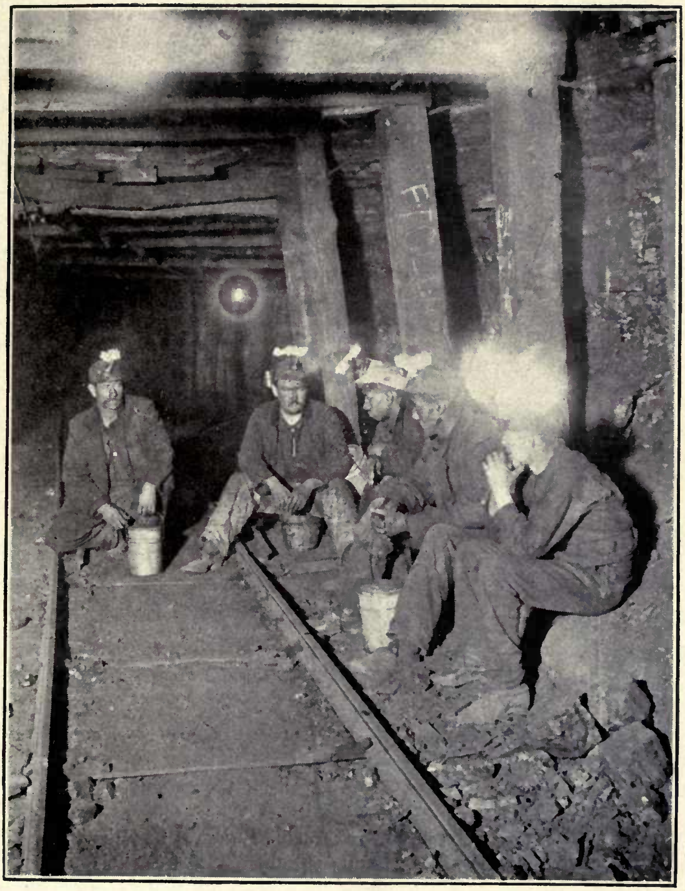
The coal is distilled in retorts made of fire-clay and about eight feet long. Six or more retorts are arranged in tiers forming a group or bench, so that all the retorts of a bench can be heated by a single fire—usually of coke.

The distillation continues from four to six hours, during which the temperature often reaches 1200° C.

The volatile products pass from each retort and bubble through water into a hydraulic main. Here some of the tar is deposited and ammonium compounds are dissolved by the water which flows constantly through the main. This water acts as a "seal" to prevent the gas from passing back into the retorts.

From the hydraulic main the gas passes into the condenser. The main object of the condenser is to cool the gas slowly and condense and remove the tar. An exhauster in most plants draws the gas from the hydraulic main through the condenser into the scrubber. The scrubber is to remove the remaining ammonia, part of the carbon dioxide and hydrogen sulphide gas, and the last traces of tar. One form is a double tower filled with coke or pebbles over which the ammoniacal liquor slowly trickles in the first part and pure water in the second. The gas enters at the bottom, meets the descending liquid, and is thoroughly washed. From the scrubber the gas passes into the "purifiers." Their chief purpose is to remove the remaining carbon dioxide and sulphur compounds. The purified gas passes through a large meter, which records its volume, into a "gas holder."

This little introduction will show that it is at present impossible to enter into any thorough chemical consideration of coal. We introduce the mention of it in this place only to bring in some consideration of the element carbon on account of its importance in all that concerns human life. You will be prepared already to understand something of the reason for the value of coal as a fuel. Carbon, combining readily with oxygen, yet combines slowly enough to keep up the chemical action that gives us warmth, steadily, effectively, and without attention. Illuminating gas was for a long time derived exclusively from coal, being distilled from it, and sent where needed. The combination of heat in so portable and effective a form, with a light so easily managed, almost recreated industries. By means of light the hours of labor were lengthened, and the new fuel meant an enormous multiplication of man's power. And this whole change, you must mark, was due to chemistry. Coal was no more than a convenient, cheap, and readily obtained form of the element carbon; and this element, together with the ones we have already spoken of, hydrogen, oxygen, and nitrogen, make up a great part of all animal and vegetable sub-



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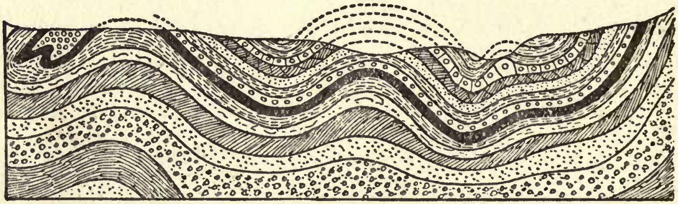
DINNER TWO AND ONE-HALF MILES UNDERGROUND.

stances. These, then, are the four great elements, and a knowledge of these four forms the basis upon which practical chemistry must ever be built up.

We have already seen how, by means of burning wood without allowing too much oxygen to reach it, charcoal—another less pure form of carbon—is produced. This will help us to understand how the light, soft pollen, dropping for centuries to the ground in undisturbed layers, and being then carried underground or overlaid by rocks and new soil, which must have been deposited by earthquakes, by seas, or by other agencies at a later time, was subjected to pressure that consolidated the powdery substances into a solid mass, and at the same time, owing to the heat produced by pressure, to some extent oxidized these layers of vegetable substance and separated them from such elements as the heat would drive off. The result was what may be called a slow oxidizing, or burning. Owing to the lack of oxygen beneath the earth, this burning acted much as the slow burning of the charcoal-maker's fire, and produced likewise a sort of rock charcoal—what we know as coal.

It is now believed that compounds of carbon are subject to precisely the same laws that govern all others. Though we still speak of "organic" compounds, we do so in a different sense. The term no longer means compounds that cannot be met with except in some living organism, but simply denotes such as are usually found there. As it happens, carbon enters into these, forming a vast number of them which are related to one another, and therefore are usually studied together. When it is remembered that these organic compounds include all vegetable products, such as sugar, starch, alcohol, honey; and all those animal products that are used in industries, such as horn, leather, meat, wax—it will be seen that practical life depends largely upon them. Though they are so various in appearance and qualities, yet they contain for the most part only a few elements, the four we have already mentioned being the chief.

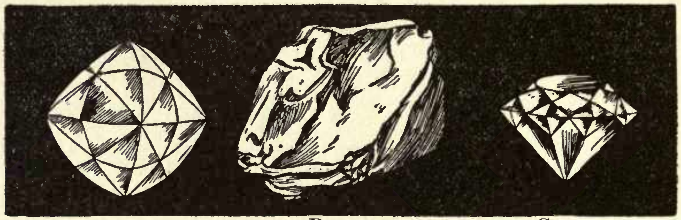


SECTION OF PART OF EARTH'S CRUST NEAR MAUCH CHUNK, PENN., SHOWING LAYERS OF COAL

To take a few examples, starch and fruit juices contain carbon, hydrogen, and oxygen. Meat contains, in addition to these three, nitrogen, and other animal substances may contain also sulphur or phosphorus. When chemists make up new compounds, such as dye-stuffs or extracts for various purposes, these may contain almost any of all the elements, depending upon the properties desired. As we know something already of the other elements mentioned, we may give some attention here to the characteristics of carbon.

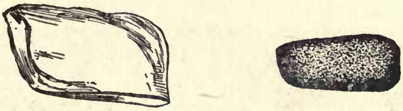
It is found pure in nature, as in the diamond and graphite, and nearly pure in charcoal or coal. In combination, we have already said it is nearly universal, being an essential part of wood, cotton, wax, starch, albumen (or white of egg), bones of all animals, and also many minerals, among which may be mentioned limestone, chalk, and marble. The purest carbon, the diamond, will give us some idea of the qualities of the element. It is the hardest of all substances; weighs three and a half times as much as the same bulk of water; no liquid at any ordinary temperature will dissolve it; yet, in spite of its hardness, it is brittle, and may be readily shattered. The proof that the diamond is pure carbon consists in burning the stone in pure oxygen. After a diamond has been entirely consumed, there will remain a gas, which analysis shows to be carbon dioxide—a combination of carbon atoms with oxygen atoms.

A still stronger proof was given by a French chemist, Moissan. This chemist dissolved pure charcoal, or carbon, in melted iron. This solution was then poured into water, which cooled the surface of the mass of iron, caused it suddenly to contract, and thus brought an enormous pressure to bear upon the carbon dissolved in the iron. After the iron was cut apart, crystallized carbon was found inside, and tests of every sort showed that these crystals were diamonds, that is, were pure carbon.



Crystal **Rough** **Cut**
Diamonds

It has already been said that carbon occurs in three forms, as diamond, graphite, and in the less pure forms found in coal, charcoal, soot, and other such charcoal-like remains of combustion. The diamond and the graphite are carbon crystallized, which means that they have been formed by carbon dissolved or melted and then allowed to cool, in which case it takes regular forms, as do such substances as sugar, salt, and alum. With graphite the ordinary lead-pencil has made us familiar.



Artificial Diamonds (Enlarged) Prepared by Moissan.

While the diamond is the hardest of all things,* graphite is soft enough to leave a mark even on paper or the fingers. It is much lighter than diamond, and is very useful in the arts because of its great resistance to heat. It can be dissolved like other forms of carbon in melted metals.

Of coal we have already spoken, but we may add to what has been said that it occurs in three or more widely differing varieties. When hardest, it is known as anthracite; a softer sort is the bituminous, which is less pure than the anthracite, containing more substances, such as oils, that are not pure carbon. A still softer sort of coal is lignite, which appears to be formed of vegetable substance that has not been reduced so nearly to carbon alone. Peat is even softer than lignite, and has not really been formed into coal, though it may be an earlier

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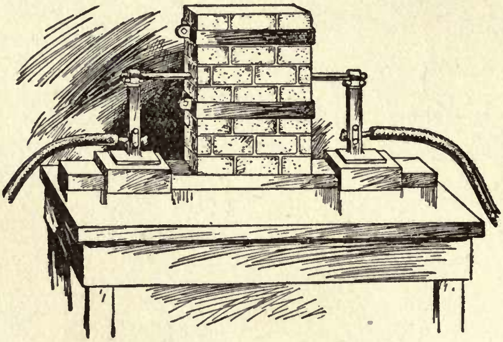
* Carbide of boron is mentioned as a possible exception in Smith's "General Inorganic Chemistry."

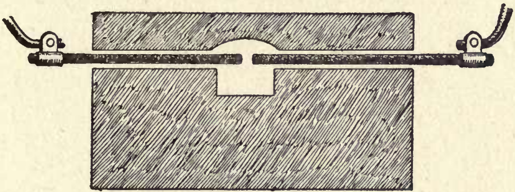


Moissan's Electric Furnace for Making Diamonds

Moissan's description of this furnace is as follows: "It consisted of two bricks of quicklime placed one on top of the other. The lower brick contained a long longitudinal groove to receive the two electrodes [carbon rods], and situated in the centre was a small cavity. This cavity might vary in size, and contained a bed some centimeters in depth of the substance to be acted upon by the heat of the arc, or a small crucible of carbon containing the substance to be treated may be placed there. The upper brick was slightly hollowed out in the part just above the arc. As the intense heat of the current soon melted the surface of the lime, giving it, at the same time, a beautiful polish, a dome was obtained in this way which reflected all the heat onto the small cavity which contained the crucible."

stage of the same process that, continued, would result in coal formation.

Ordinary charcoal may be produced from almost any kind of animal or vegetable matter that is partly consumed, that is, combined with oxygen so as to leave a large part of the carbon unchanged. If soft coal be burned much as wood is burned in making charcoal, the substance that remains is known as coke, which is about nine-tenths carbon, but harder and heavier than ordinary wood-charcoal. It is used in many industries, since it burns with a smokeless, intense heat, and contains little sulphur or other volatile substances that are harmful in such processes as making iron.



Vertical Section of Moissan's Electric Furnace. EE are the Electrodes

This shows the furnace with the parts slightly separated. The furnace is small, some being only 16 to 18 cm. (about 7 in.) long, 15 cm. wide, and 14 cm. high. The carbon rods are from 1 to 5 cm. in diameter.

The three main forms of carbon are, nevertheless, all the same substance. The difference in their form and quality is believed to be due to the fact that the molecules in each case contain a different number of atoms. This character of assuming different forms without a change of substance is known in chemistry as allotropism, from *allos,* another, and *tropos,* a form. Sulphur is another element that shows allotropism.

Owing to the fact that carbon is so widely found and forms compounds that exist in all living things —plants and animals—a later chapter must take up the methods in which its compounds are formed. But in an elementary book only the most general account of carbon's manifold compounds can be given.

CHAPTER VII

NATURE OF CHEMICAL COMBINATIONS

WE have learned that burning is to be understood as a combining of oxygen with some other element. There is no attempt to explain why such combining takes place. We only know that once certain elements are raised to a certain degree of heat they begin to form new compounds with the oxygen, giving rise to substances of entirely different properties from those that have been burned.

But this is not the only example of such compounds. Things can be burned in other gases than oxygen. Chlorine will thus combine with hydrogen and form an acid known as hydrochloric acid. Both copper and iron may be caused to unite with the vapor of sulphur, giving copper sulphide, or iron sulphide. Other chemical combinations are formed by simply bringing together solutions, or solids and solutions, of different elements. Thus if we put pieces of zinc into hydrochloric acid we will find that the zinc will drive the hydrogen from its compound with the chlorine, forming a chloride of zinc, while the hydrogen is set free.

All these changes depend upon qualities that are known as "chemical affinity," which is another way of saying, chemical liking. The reasons for this affinity are not known with any certainty. We can only say that under some circumstances a substance has a stronger attraction for one substance than for another, and thus when compounds are brought together there is a change of partners, as it were. The changes in compounds thus brought about are known as double decomposition, for the simple reason that *two* compounds when brought together are *both* separated and then new compounds are formed from the partners. Thus, if oxygen and mercury exist in one compound, and chlorine and hydrogen in another, by bringing them together they exchange partners, the hydrogen and oxygen uniting to form water, while the mercury and chlorine also join forces, making mercury chloride.

It is also true that certain compounds are strongly joined and are difficult to separate into their elements, and that others seem to be joined by a weaker bond and readily separate. Thus it is enough to heat certain compounds slightly to bring about a change in their composition. Others will not be separated until we have found some element having for one of the partners an attraction exceeding that for its partner in the first compound.

Thus it happens that the chemist in bringing about changes is not always certain beforehand just what will take place. Sometimes by knowing which elements strongly attract others, he can foretell just what new compounds will be made. In other cases he cannot be certain until the experiment has been tried.

Two elements which seem to have an exceedingly strong attraction for one another are hydrogen and oxygen. Consequently whenever compounds containing hydrogen and oxygen are brought together it is very likely that the hydrogen and oxygen atoms will join, forming water. Understanding this, we may see the reason for a very common result in the mixing of two compounds in liquid form. If one contains hydrogen and the other oxygen, and each contains also two substances that together form a solid, a mixture of the two first compounds will result in combining the hydrogen and oxygen, whereupon the remaining partners will be left to unite into a solid, and the result of the experiment will form a solid (in fine particles) floating, or diffused, in water. The chemical change being complete, we may then allow the mixture to settle, or may filter it, and in this way the water may be poured off, leaving the solid. If the solid is one that will not float and is allowed to settle, it is known as a precipitate.

It will surprise not even the youngest reader to be told that there is practically no limit to the number of chemical compounds. There are seventy or eighty elements, and while a few of these are not ready to make combinations, and a number of them are very rare, at least twenty of them are very important, and many of these are exceedingly common. If the reader will reflect upon the number of words that may be composed out of the twenty-six letters of the alphabet, he will gain some idea of the almost infinite number of possible chemical compounds, to say nothing of the mixtures of these or of their different states, as solids, liquids, and gases. Fortunately, however, the larger part of these compounds may be generally classed under a few headings, and the more important of these agree in certain general characters. The names of these classes are more or less familiar to us all from daily use. Thus we know that chemicals are divided into acids, bases, and salts. Although these three do not include all chemical compounds, they are convenient names under which to group the larger part. The acids, for example, were so called because in this sort of compound the early chemists detected a sour taste. Acids also changed the color of many vegetable substances, and are for the most part active compounds ready to enter into chemical changes. The bases are such compounds as when acted upon by acids form salts. Their character will be explained later.

The acids all contain hydrogen, and most of them contain also oxygen. It was at first believed that oxygen was the maker of acids, and the gas was named from this belief. But it has been found out that hydrogen is the essential element in acids. Joined with the hydrogen (or hydrogen and oxygen), to make an acid, is always some other element among those known as the non-metals, examples of which are sulphur, chlorine, phosphorus, and nitrogen. The bases contain oxygen, usually hydrogen, and always a metal; and elements which form bases are ranked with the metals. The salts contain a metal, a non-metal, and most of them also oxygen.

It is the chemical relation of these three classes of compounds that really fix their classification. Thus when an acid and a base are made to act upon one another, they destroy more or less each other's properties, and the result is a compound having few or none of the properties like those of the two that formed it. This resulting compound is a salt. For example, if we add hydrochloric acid to sodium hydroxide, the result may be written in the following formula:

$$HCl + NaOH = NaCl + H_2O$$

the meaning of which is that hydrochloric acid is neutralized by the sodium hydroxide, and that the result is sodium chloride (common salt) and water. In this case we have an acid, combined with a base ($NaOH$), and the result is a salt. The metal, sodium, has taken the place of the hydrogen in the acid, while the hydrogen of the acid is combined with the hydrogen and oxygen of the base, forming water. Something more of this action is told later, in speaking of electrolysis and solutions.

Let us now look a little more closely into the nature of acids, bases, and salts.

Although the definition of an acid is not an easy thing to write in simple terms, yet the nature of an acid may be readily understood from a few simple examples. An acid in general is strong in its action, attacking—that is, changing—the composition of compounds, or taking into its own composition atoms of other elements. The commoner acids are sulphuric, nitric, and acetic. The acids themselves occur as solids, liquids, and gases, but most of them are soluble in water, and, ordinarily speaking, it is this solution that we mean by an "acid." When the acid is very strong, or mixed with little water, it must be handled with the greatest care, since its action is rapid and destructive. When much diluted, the action is of course slower and the other qualities are likewise diminished, and the solution may be more safely handled.

The way in which acids are ordinarily recognized in the laboratory is by means of a test with what is known as "litmus" paper. This is a paper containing vegetable juices procured from a species of lichen. This juice stains the paper a soft blue color, but when dipped into, or touched by, the acid, the blue is changed to a pink or reddish tinge.

Having been so changed, it may be restored to its original color by being dipped into an alkali. This is a delicate test and a very convenient one. It may be familiar to young readers, since it is used in photographic operations to determine whether certain baths, as those used in toning prints, are acid or alkaline.

The sour taste of acids is, when not too strong, very agreeable to the palate, and has led to the use of many substances for food because of the acid they contain. Vinegar is water flavored with acetic acid; pickles contain the same acid; lemon juice and other fruit juices are relished because of the fruit acid in them. Milk, when it turns sour, does so because it forms an acid—lactic acid.

With these few examples, let us see what the chemists say in the attempt to define an acid in chemical terms. The books tell us that an acid is "a compound containing hydrogen that can be replaced by a metal." The meaning of this is that when a compound so acts upon a metal as to give up hydrogen and to replace this by atoms separated from the metal, this compound is to be classed with the acids. But although no one ignorant of chemistry would think of calling water an acid, under this definition it must be so called; for water, in acting upon certain metals, will give up hydrogen and produce an oxide of those metals. For this reason chemists agree that, rightly speaking, water should be regarded as an acid, although it does not affect the color of litmus paper.

In the early days of chemistry it was believed that the important element in all acids was oxygen, but as time went on and further experiments were made and understood, compounds were found that certainly must be considered as acids, and yet were entirely free from oxygen. Hydrochloric acid (HCl) is an example. It is for this reason that the definition given above, making hydrogen the chief element in the defining of acids, was adopted.

Nevertheless, it is true that most acids do contain oxygen, and the ordinary names of the varieties of acids and of the substances resulting from their action are based upon the *amount of oxygen* contained in the acids. To that form of an acid which is best known is given a name ending with the letters *ic*. Thus an acid made up of sulphur, hydrogen, and oxygen (H_2SO_4), is known as *sulphuric acid.* But another acid can be formed from these same elements containing less oxygen, so the ending of its name changes, and it is called *sulphurous* acid. If a third form has still less oxygen, it is named by putting *hypo* before the names ending in *ous*, as, hyposulphurous acid. This word, hypo, is from the Greek, and means *under*.

In the same way we may form the names nitric and nitrous acid, or chloric, chlorous, and hypochlorous. If, however, there be a form of any acid containing even more oxygen than the commonest and best known form, this is expressed by the prefix *per* added to the *ic* form. In this way we have persulphuric acid and perchloric acid.

These can be remembered by simply recalling that *per* means more, *hypo* means less, *ic* is attached to the commonest acid, and *ous* to one that is weaker in oxygen.

Owing to the peculiar chemical qualities of carbon, acids in which carbon is an important element are not named precisely according to this system, but they end ordinarily in *ic*.

These names are *chemical* names, but there exist also "trade," or commercial, names for some of these acids. A list of these common terms is given in the last chapter.

The word base is used in chemistry to mean the class of substances which neutralize acids; that is to say, which, when caused to combine with the acid, will destroy its properties. Most of these are solids, but just as we in speaking of "acids," usually have in mind the solution of these substances in water, so in speaking of bases we often mean such as may be made into solutions that will neutralize the acids. Examples of such bases are lime, ammonia, sodium, and potassium. When these are referred to with their property of neutralizing acids in mind, they are commonly called "alkalies." They are tested by means of litmus paper, and have the property of restoring to the reddened paper its blue color. Generally speaking, a base is a substance derived from a metal by the action upon it of hydrogen or oxygen, and having the property of combining with an acid to form a salt. Ammonia, however, does not contain a metal, and yet is a base, for it contains a combination of elements (NH_4) that acts as a metal. A name often applied to certain bases is "hydroxides," or hydrates. This name is given because most of them contain hydrogen and oxygen. By putting before the term, hydroxides, the name of the metal from which it is derived, we distinguish the base. Thus we have sodium hydroxide, calcium hydroxide, calcium hydrate (slaked lime), and so on.

When an acid and a base are caused to combine, each loses its properties and enters into a new compound not like either. Compounds resulting thus from the action of an acid and a base are known as salts—probably from their resemblance in many properties to common salt (which is, chemically, chloride of sodium). Most salts are soluble in water, and when such solution is made, it will be found that litmus paper may be dipped in it without changing color. Thus the acid which would have changed the paper red, and the base which would have turned it blue, have now united into a substance which has neither effect. The salt, therefore, is said to be neither acid nor alkaline, but "neutral."

Salts are not the only neutral substances; neither are all salts neutral. Nevertheless the general rule is as stated. The name salt is also applied to certain substances produced otherwise than by the action of an acid on a base, since the name is used only as a convenient way of classing certain compounds together.

The names of different salts have been given in such a way as to show which class of acid gave rise to them. The acids ending in "*ic*" form salts ending in "*ate*." Thus sulphuric acid acting on zinc would give zinc sulphate. But those ending in "ous" form "*ites*"; Sulphur*ous* acid forms sulph*ites*. If the prefix *hypo* or *per* occurs in an acid, it appears also in the name of the salt. Thus a perchloric acid would give a perchlorate; and hyposulphurous acid would give a hyposulphite—a name which will be recognized by photographers as belonging to the well-known fixing solution, which is a hyposulphite of soda. That is to say, it is formed by hyposulphurous acid acting upon the base, sodium, to make the salt, hyposulphite of soda.*

Where a salt contains only two elements, the name ends in *ide*, this being affixed to the non-metallic element. Thus the salt is known as a bromide, chloride, sulphide, and so on, depending upon the element upon which the action of the acid has taken place. Where *hydro* occurs in the name of the acid, it is omitted when the name of the salt ending in *ide* is given. Thus we have in the name for common salt, sodium chloride, the fact that it is produced by the action of hydrochloric acid upon sodium, and that it is a compound of only two elements, which we know because it ends in *ide*. The prefix *hydro* is not necessary when only two elements are in the salt, because it is known that the hydrogen or oxygen from the original acid have combined into water, since neither of them has entered into the salt. If either were in the salt it would have more than two elements. Thus to produce common salt the hydrochloric acid has separated and given its chlorine to the sodium, making chloride of sodium. This is the equation:

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* This illustration may be retained, although, strictly speaking, the fixing solution is chemically called "sodium thiosulphate." This syllable, *thio*, indicates that this special salt is the result of the action of thiosulphuric acid on sodium, and thiosulphuric acid indicates an acid containing more sulphur and less oxygen than sulphuric acid. This, however, does not affect the reason for the old name of hyposulphite of sodium.

$$HCl + Na = NaCl + H$$

Showing that one atom of hydrogen is set free, or, it may be, is taken up by oxygen to form water.

CHAPTER VIII

ABOUT COMMONER ELEMENTS

FORTUNATELY for those who wish to know a little something about chemistry without too much deep study, the elements that are commonest are not very numerous, and even with these commoner elements we can aid our memories by grouping them into certain families, or classes, the members of which resemble one another in their action and in their properties. Thus one wide grouping is caused by a division into metals and non-metals, another into the acids and the bases, and the salts.

When certain elements are united with hydrogen and oxygen to make compounds, some of them will be found to belong to the acid compounds and others to the base compounds, and the base-forming compounds include what we know as metals, and also certain other elements that are called metals because they are grouped with these. This causes the other group of elements, the acid-forming group, to be referred to as non-metals.

It is not true, however, that this distinction always holds, for a few elements may, with hydrogen and oxygen, form compounds sometimes of one kind and sometimes of another. But these matters may be mentioned in the case of each element.

For the reasons given, chemists have for a long time put the elements into groups, of which we may consider the following as principal ones: first, hydrogen, which though it may be considered a non-metal, stands alone, making, as it were, a group by itself, for no other element bears a very close relation to it. Next, oxygen, which, although not strongly related to any other element, yet has some resemblance in the character of its compounds with the element, sulphur, and therefore may be thought of with it. Next comes a group of four elements, the halogens or "salt-makers," which includes chlorine, bromine, iodine, and fluorine; then one of three, including sulphur, selenium, and tellurium. To this group oxygen is sometimes added. Another group of four is the nitrogen family: nitrogen, phosphorus, arsenic, and antimony; then follows a family of two members, carbon and silicon.

We have not named all the members of these families, but given only the commoner elements belonging to them. All these elements named belong to the larger class already spoken of as the "non-metals" (see List of Elements in last chapter), except that one of the metals, a member of the nitrogen family (antimony) also forms bases, and thus acts like a non-metal.

In order to get some idea of the character of these elements, we shall tell something of the action of each group separately, taking up first the chlorine family—for we have already spoken of hydrogen and oxygen.

THE CHLORINE FAMILY, OR HALOGENS

This family has a special name of its own. The members of it are called halogens. This term comes from the Greek words, *hals* (ἅλς) and (γενής). Hals means salt, genes means producer; so the group is known as salt-makers.

CHLORINE

Chlorine, the first of these, is a gas at the ordinary temperature. It was first prepared in 1774 by the German chemist, Scheele. Since chlorine readily combines with a great many common substances, it is not found in nature except in compounds, the commonest of which is common salt, NaCl. This is a combination of chlorine with sodium, and is believed to consist of one atom of each combined to make one molecule of sodium chloride. Chlorine can be separated from common salt by treating it with a compound known as manganese dioxide, sulphuric acid, and water. The experiment is rather complicated. When the chlorine gas is separated, it is seen to be a greenish-yellow color. This gives it its name, from the Greek, chloros, which means "green"; and its odor, in small quantities, resembles that of seaweed. In large quantities it is dangerous to inhale it, as it causes a severe inflammation. It is more than twice as heavy as air, and when it is put under pressure about five times that of the atmosphere (80 pounds to a square inch), it condenses to a yellow liquid.

The chief use of chlorine in manufactures is as a bleaching agent. It is to chlorine that chloride of lime owes its power as a bleacher. It is also used as a disinfectant and deodorizer and as a solvent of gold. Gold ores are crushed, and the chlorine unites with the metal to form gold chloride ($AuCl_3$), from which the chlorine may be separated by the use of other chemicals. Its action in disinfecting and bleaching depends upon the fact that chlorine readily combines with hydrogen, and thus separates it from water (H_2O), leaving the oxygen free. The oxygen, when just set free, is believed to be separated *into its atoms,* and these atoms thus set free, and called "nascent," combine with other substances, even when the same atoms in a molecule of oxygen would not have this power. When combined with hydrogen, the chlorine forms hydrogen chloride, or hydrochloric acid. It is formed by bringing together equal volumes of the gases, hydrogen, and chlorine. These, when exposed to the light, at once join, making the same volume of hydrochloric acid gas. If the light be strong, the two gases unite with a violent explosion. The symbol for hydrochloric acid is HCl. As already stated, it is also called muriatic acid.

There is a process for obtaining chlorine from hydrochloric acid gas, by heating it with oxygen, whereupon the oxygen unites with the hydrogen to form water, leaving the chlorine free. Another method is to decompose common salt by electricity, much as we have seen that water can be separated into its gases by the same agent.

Since it is so much heavier than air, chlorine gas may be allowed to flow into a jar of air, whereupon it collects at the bottom and gradually forces the air up and out.

Just as charcoal and iron may be burned in pure oxygen, certain elements will burn in chlorine. Thus antimony and arsenic sprinkled into it combine with a flame. Phosphorus it first melts and then burns. Several metals, when heated, will continue to burn in chlorine, but chlorine does not combine with carbon directly, and therefore a piece of glowing charcoal thrust into it is put out. The atomic weight of chlorine is 35.18.

BROMINE

Bromine (Br) occurs in nature in common with chlorine, and is likewise found combined with sodium. Its atomic weight is 80.* Bromine is prepared in two ways. It is extracted from potassium bromide by using sulphuric acid and manganese dioxide; or by passing chlorine into a solution of potassium bromide, the bromine is replaced by the chlorine and so set free. Electricity also may be used to separate bromine from bromides that are soluble. It is a heavy, dark red liquid, easily changed to a brownish red vapor. It gets its name from its bad smell, bromos, in Greek, meaning a stench. It acts much like chlorine, combining with many elements very readily. It is somewhat weaker than chlorine, and chlorine brought into contact with its compounds will often replace it. It was discovered in 1826, in sea water, by a chemist named Balard. It is very volatile, easily giving off fumes. It has a destructive effect on the skin, burning it terribly.

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* For correct atomic weights, see the tables. In the text round numbers are often used.

It is principally used as a disinfectant and in preparing dyes. It is a bleaching agent, like chlorine, though weaker. In medicine, potassium bromide, or bromide of potassium (KBr), is used as a sedative. The bromide of silver is of great importance in photography, being used to make printing papers and sensitive emulsions.

IODINE

Iodine (I), the atomic weight of which is 125.9, is a solid, dark bluish-black, with a metallic lustre. When heated slightly it is changed to a beautiful violet gas, and from this gets its name, the Greek name meaning violet. It is less irritating in vapor than chlorine, and has feebler bleaching powers.

Like chlorine and bromine, iodine is used as a deodorizer and disinfectant. It is prepared in much the same way as these two elements. Its vapor is nearly nine times heavier than air. It dissolves slightly in water and freely in alcohol, ether, and several other solutions. It combines so readily with phosphorus that when brought together the phosphorus burns. Iodides, as will be seen, are in their general action much like the chlorides and bromides. It is much used in medicine, to prevent the spread of eruptions, to reduce swellings, and to dress wounds.

FLUORINE

The remaining member of the family is fluorine (F). Its atomic weight is 19. It is a greenish-colored gas, with a strong, irritating odor, and readily attacks most substances, that is, forms new compounds with other elements. It is of recent discovery, and only important because of one or two of its compounds. Most important of these is hydrofluoric acid, HF. This is a very active, colorless gas, valuable in the arts because it readily dissolves, or corrodes, glass, and so may be used for etching patterns on it. This is done by covering glass with a thin coating of wax, then removing the wax in a pattern, and exposing the whole to the vapor of the hydrofluoric acid. These fumes eat into the exposed glass, but do not attack the wax. Warm water then will remove the wax.

Its density is 1.265, very slightly greater than air. Its activity is shown by the fact that hydrogen, bromine, iodine, sulphur, phosphorus, carbon, silicon, and boron all take fire when immersed in it. Most metals form fluorides with it. It does not unite with oxygen. But fluorine is the only element that does not combine with oxygen.

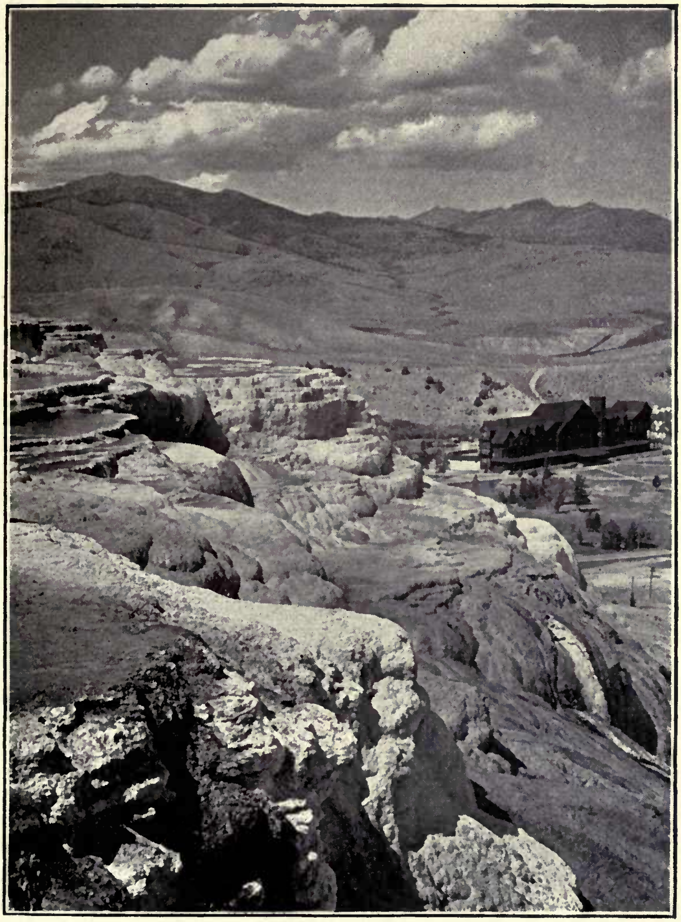
This completes the account of the fluorine group, and the close relation between its members is readily seen. Thus, all form colored vapors. They combine strongly with many substances, fluorine having a stronger affinity, when it combines at all, while chlorine, bromine, iodine, follow in that order. All of them combine in single atoms, and when combined with other elements, compounds made in the same way resemble one another in properties. These properties seem to depend in some way upon the atomic weights of these elements, as the properties change in the same order with the increase of atomic weights, as will be seen later.

It will be found that the same thing is true in regard to other chemicals having a similar family relation. Thus, if we know the atomic weight of an element, we can, to some extent, foretell its properties, though the reason for this is not yet known. Putting down in order the weights of fluorine, chlorine, bromine, and iodine, we shall have 19, 35.5, 80, and 127. Roughly, these numbers show a regular progression in the order of the strength of their action.

The second family is the sulphur family, consisting of the very common element, sulphur, and the rarer elements selenium and tellurium, and in connection with them oxygen is often grouped since it somewhat resembles them. Selenium and tellurium are exceedingly rare, but form compounds—selenides and tellurides—resembling those of sulphur.

THE SULPHUR GROUP

There are four elements ranged under this head, and one of them is oxygen, concerning which much has been said already, and of the other three the only important one is sulphur, the two others being very rare elements of very little importance except to professional chemists, and having no characters



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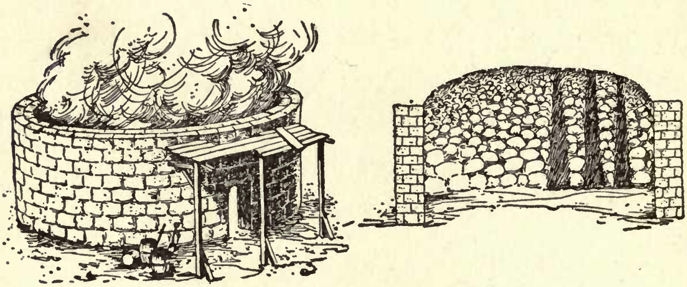
SULPHUR SPRINGS.

we need remember, except that their compounds, in general, are not very unlike those of sulphur.

SULPHUR

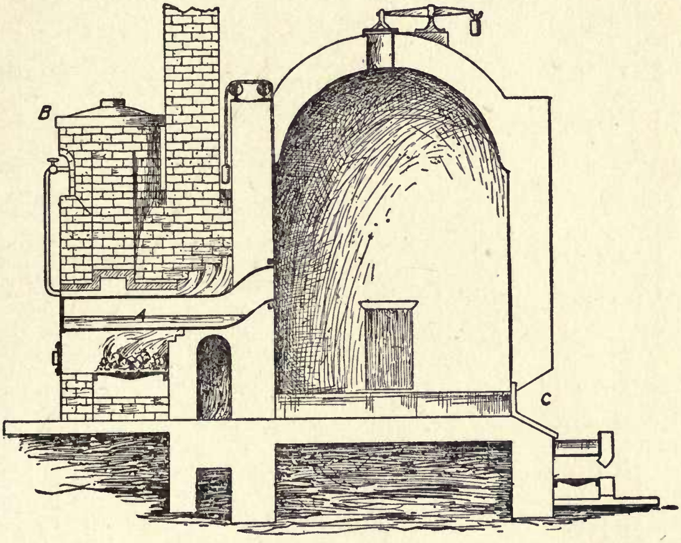
Sulphur itself, however, is exceedingly important. In the first place, it has been known since the very earliest times, probably because it occurs in large quantities in its pure state in volcanic regions; secondly, because it is so peculiar in appearance and has such striking properties. It is also found in a pure state, or nearly so, near beds of gypsum, which itself is a sulphur compound. A great deal of it, nowadays, comes from Sicily and from South America. It exists among different rocks in its various compounds, and is also found in many articles of daily food, being important especially in onions, horseradish, garlic, and in the yolks of eggs. It enters into the composition of animal bodies, of which it forms a most important constituent. In the human body there is a little more than a quarter of a pound of sulphur in various forms. It is produced, when not found pure, by heating what is known as its "ore," which is purified and the sulphur melted from it, or washed from it by hot water or steam. After it is liquefied, still further heating makes it a vapor, and this vapor, when condensed on a cool surface, takes on a powdery form.

In its commonest form it is yellow, brittle, and often in crystals. It weighs about twice as much as water. When heated, it at first melts into a light amber-colored liquid, then gradually darkens and thickens into a very sticky, dense liquid. Further heating causes it to thin again, finally to boil and change to a vapor. If it catches fire, it burns with a pale blue flame. It combines readily with hydrogen, carbon, chlorine, and most metals, making compounds of two elements, or "sulphides." Chemists place it among the non-metals, and find that it exists in a number of forms, or is allotropic.



KILN FOR EXTRACTING SULPHUR FORM CRUDE ORE
It is shown as a vertical section (right) and in operation (left).

With lead it forms the lead sulphide, "galena"; with mercury, it forms the sulphide "cinnabar." Both of these ores are easily freed from sulphur, giving the lead and mercury pure. Cinnabar is useful to artists, being the pigment vermilion, though this color is the very pure sulphide artificially produced.



Apparatus for Purifying Sulphur

The crude sulphur is melted in B, and flows into the iron cylinder, A. Here it is heated, and the vapors pass into the large brick chamber, provided with a tap, C, from which the liquid sulphur may be withdrawn.

A mass of sulphur, when rubbed, becomes electrified negatively, and a ball of sulphur mounted on an axis so as to be revolved against a rubber, made the first electric machine.

There are at least six different sorts of sulphur known to chemists, though we shall not try to describe them very accurately.

1. In crystals known as orthorhombic; that is to say, crystals having three unequal axes crossing at right angles. This seems to be the primary form of the sulphur crystals, for other forms, left to themselves, gradually change back to this one.

This form of sulphur is soluble in ether, benzine, and some other chemicals, but not in water.

2. When sulphur is first melted and allowed to cool, it forms what are called *monoclinic* crystals, or crystals with three unequal axes, two crossing obliquely, and the third at right angles to the plane of these two. If this form of sulphur be left to itself, it will, as has been said, gradually change to the first form. This is soluble in alcohol and in ether, and in some other liquids, but not in water.

3. A soft, formless, whitish form of sulphur, believed to contain some hydrogen. This, when left

to itself, gradually hardens and changes to the first form.

4. An insoluble form of sulphur, that is plastic, or may be moulded.

5. A form called "amorphous" or formless, yellow, and insoluble.

6. A form known as "colloidal."

These last three forms are distinguished from one another by their differences in solubility, and so forth. The different forms also do not agree in specific gravity and some other qualities.

As ordinarily sold, sulphur comes in long sticks, or rolls, known as brimstone, or rolled sulphur. These, crushed, become ordinary sulphur. Very finely powdered sulphur is known as flowers of sulphur. Sulphur is very useful medically.

Its symbol is S; its atomic weight 31.82. The importance of knowing its different forms depends upon the fact that these help us to make guesses at the arrangement of its atoms in making up the molecules. Since the material composition, or make-up, of an element is always the same, when we see that certain changes in its condition cause it to assume, as in the case of sulphur, six varying forms, we can only explain these by supposing that the little atoms forming its molecules are arranged in some different manner or relation. And this becomes very important when we study the compounds of the various elements, especially those of carbon. Many of its compounds can only be understood by supposing that the same atoms are arranged in different ways in their make-up.

The important compounds of sulphur are many. Thus, with hydrogen it forms H_2S, or sulphuretted hydrogen. Its chemical name is hydrogen sulphide. It is a gas, occurring in the waters of sulphur springs and in volcanic gases. It is one of the impurities of illuminating gas, and may be found in the air near sewers and cesspools, being sometimes formed by the decay of organic substances which contain sulphur; thus it is formed when eggs decay—and causes their bad odor. It is colorless and poisonous. It is useful in the chemical laboratory as a "reducer," that is to say, as a means of extracting oxygen from compounds. If a house be heated with coal, hydrogen sulphide in the air will often cause silverware to tarnish, the tarnishing being due to the formation of silver sulphide. A similar chemical action is caused upon silver spoons when these are brought into contact with mustard or the yolks of eggs. White-lead paint, whether upon the walls of a house or on oil-paintings, is also tarnished by this gas, becoming greyer. Hydrogen sulphide is often useful in chemical analysis, as when sulphides of metals are formed by its use, they may be recognized by their different colors.

SULPHUR DIOXIDE

Sulphur dioxide is the common compound of sulphur with oxygen, which is important particularly in the manufacture of sulphuric acid. Sulphur dioxide is also used for preserving food, for fumigation, in tanning, and in the refining of sugar, and in a dozen other ways.

SULPHURIC ACID

But the most important of the sulphur compounds is sulphuric acid. The formula for this is H_2SO_4. In an article on the part played by chemistry in human progress, Professor Austin, of Rutgers College, says: "If one were to ask how the state of civilization of a country should be judged, one would probably be told how many churches there were, attention would be drawn to the volume of business done, fine hospitals, the excellence of the police," and so on. " But," he says, " the chemist would simply determine the amount of sulphuric acid used, directly or indirectly," for he declares that nearly everything used by civilized people requires in its making this acid.

The uses of this acid are beyond enumeration, and it has been well said that the health and prosperity of a nation can be measured by the amount of it in use. A large part goes to the making of fertilizers. It is used in making all other mineral acids, and many others besides. It is necessary in one process for making sodium carbonate, a compound that enters into all soap and glass. It is used in the manufacture of alum, nitro-glycerine, glucose, phosphorus, dye-stuffs, in bleaching, in electric batteries, refining and metal working, and, in short, is the commonest of the acids used in the arts. In fact, there are very few of our industries which do not depend more or less upon sulphur or some of its compounds, to say nothing of its direct uses, such as in making gunpowder and fireworks, in the making of matches and in the hardening of crude rubber. We are so used to what we know as hard and soft rubber that we do not realize how recent is the discovery of a means for making this useful substance. It was only after years of study, and to some extent by happy accident, that Charles Goodyear discovered the means of hardening rubber by adding sulphur to it—that is, of "vulcanizing" it.

Sulphuric acid is what is known as di-basic, that is, it forms two classes of salts, the normal and the acid salts. The sulphates thus formed, either naturally or artificially, are most important. The sulphate of calcium is gypsum, known in various forms as plaster of Paris and alabaster. Zinc sulphate makes white vitriol, while copper and iron form respectively blue and green vitriol. Sodium sulphate is the well-known Glaubers salt. Magnesium sulphate is Epsom salts. Sodium thio-sulphate is, as has already been explained, the correct name for what we call hypo-sulphite of soda, the fixing salt of photographers. With carbon, sulphur forms carbon di-sulphide, a liquid that is very useful from its property of dissolving rubber, various gums and fats, and so on. Rubber dissolved with it makes rubber cement.

SELENIUM

Selenium, whose symbol is Se, and atomic weight 78.6, is a non-metallic element occurring in two forms (1), a red powder, and (2) a dark glassy mass sometimes crystallized. This element was discovered by Berzelius, in 1817, and has no other interest for the general reader than a resemblance in its compounds to the similar compounds of sulphur. Selenium is important to the electrician because its resistance to electricity varies with the amount of light. This property is used in the "photophone," where a ray of light transmits speech by aid of a selenium cell; and in other ways.

TELLURIUM

Tellurium, symbol Te, with an atomic weight of 127.6, was discovered in 1782 by von Reichenstein. It is a silver-white metal, brittle, and also forms compounds generally resembling those of the other elements of this group.

Simply to complete the group, we name here also oxygen, the gaseous non-metallic element, whose atomic weight is 15.88. Its properties have already been discussed. But while considering it in this place, it may be well to say that it is sometimes taken to have an atomic weight of *exactly* 16. You will remember that the meaning of atomic weight is simply weight compared with an atom of hydrogen. Now if hydrogen atoms be taken to weigh 1, the oxygen atoms must be taken to weigh 15.88. But if, on the contrary, we choose oxygen as being the measure and call it 16 exactly, then a very slight change will be necessary to adjust all other atomic weights to this new scale. Hydrogen, instead of being 1, will now become 1.008, and as we go through the list of elements giving the atomic weights the slightly larger values that is made necessary by considering oxygen to weigh 16 instead of 15.88, we shall find that a large number of elements change their values from fractions to whole numbers. Among these are arsenic, boron, carbon, fluorine, manganese, mercury, radium, and tin. As it is more convenient to work sums in the whole numbers, chemists are now beginning to adopt these new values, merely as a matter of convenience; and even where a table is given showing one set of values (with hydrogen 1) other values with oxygen equals 16 are usually written after them. The change is very slight, and very convenient. The table of elements in the last chapter of this book gives both atomic weights.

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CHAPTER IX

THE METALLIC ELEMENTS

FOR ages chemists believed that there was a sharp division between the classes, metals and non-metals. Those elements which were opaque, shiny, heavy, hard, and malleable, were called metals; those elements which did not have these properties, the non-metals. Some elements, such as antimony and bismuth in certain ways appeared to be metals and acted like them; in other ways they seemed to belong to the non-metal class. They are therefore still referred to as "metalloids," or metal-like elements, though strictly speaking, they are metals.

Still, though the definition of metal and non-metal has now been changed to express chemical rather than physical qualities, yet the distinction is a useful one, and need not cause confusion. In general, therefore, we may say of metals that they are shiny, and, when pure, that they do not (excepting the thinnest silver and gold foils) transmit light. Many metals are white, such as silver, sodium, aluminum, mercury, magnesium, and iron. Iron, of course, soon takes on a red surface if exposed to the air, because oxide of iron is red. Bismuth is reddish white; gold is yellow; copper is the only metal element that is red. The quality of being malleable and ductile belongs to most metals, though in different degrees. Hardness varies very greatly, from the liquid mercury or such soft elements as sodium and lead which are readily dented or cut, to such hard metals as iridium, which is as hard as steel, or harder. The specific gravity of metals, or their weight compared to water, also varies widely.

In the last chapter, among the tables, we give the physical qualities of the best known metals.

METALS

It is characteristic of metals that they are specially good conductors of electricity and of heat. Chemically, they unite with oxygen to form oxides, and with hydrogen and oxygen they make bases, whereas the non-metallic elements form compounds that produce acids.

In nature metals occur but seldom in a pure form, the only two at all commonly so found being gold and copper. Gold is found pure usually in very small amounts, in mere grains, though exceptionally now and then immense nuggets have been found in certain localities. Copper, along our Great Lakes, occurs sometimes in large masses, almost pure.

But, for the most part, owing to the fact that the metals combine readily with other elements, they are found in their ores, or compounds, most often with one of the four common elements, carbon, hydrogen, oxygen, and sulphur. In their ores, silver, tin, iron, and lead, are not only plentiful, but are found in many districts throughout the world. The pure metals can be obtained from the ores by various methods, the treatment including especially the use of heat, electricity, and a wide variety of chemical processes.

The mixtures of metals, whether chemically united or not, are usually called "alloys," except in the case of mercury. Mercury mixtures, which are in consistency somewhat between the liquid state of mercury and the solid state of the metal mixed with it, are called "amalgams."

Since alloys may be made without true chemical union, they differ in composition in every degree.

The more important of them will be found mentioned under the metals of which they consist. They are mixed from the melted metals, and vary much from their constituents, as in color, hardness, conducting power for heat or electricity, and especially ease of melting. Sometimes an alloy will melt at a far lower degree of heat than any of its constituents. Many metals will mix readily and in widely varying quantities, while others do not unite readily.

The properties of metals often vary widely, according to how the metals have been prepared.

GOLD

In discussing the elements, we shall begin with gold. This has been known to mankind since the earliest times, and has been highly valued primarily because of its great beauty. It has an attractive color, a bright yellow that is unchangeable under all ordinary treatment. It is so very malleable and ductile that it lends itself readily to use for ornaments of every kind. Thus it can be beaten into flat plates, or gold-leaf, so thin that the smallest quantity may be extended to cover broad surfaces. It will draw into wire of the finest thinness without breaking, and in these forms, leaf and wire, may be so disposed as to make a great show with a very small amount of material. In the earliest times the fact that it was valued by every one gave it an exchange power that soon made it perform the offices of money. Being small in bulk, ready to use with little treatment, and adaptable to so many uses, all nations and all individuals were glad to get it, and it thus offered a convenient means of keeping wealth on hand. As knowledge of the arts advanced, it became usual to mix gold with a proportion of copper or silver, to prevent its wearing away so easily, and this led to the adoption of a standard expressing the proportion of pure gold in the mixture. The ordinary means of measuring this is by carats, twenty-four carats being considered the fineness of pure gold. If two parts in twenty-four are made up of another metal, the gold is considered to be twenty-two carats fine. Jewelers' gold, for the making of ornaments, ranks thus from twelve to twenty-two carats, anything lower than twelve containing too little gold to be acceptable, and anything higher than twenty-two wearing away too readily.

Chemically, the symbol for gold is derived from the Latin word *aurum,* being Au. Its atomic weight is 197, its specific gravity 19—making it by far the heaviest of the commoner metals, only platinum being heavier. Gold is less easily melted than silver, more so than copper. Pure gold is harder than lead, but softer than silver. A single grain of gold can be beaten out into a leaf covering an area of eighty square inches, and the metal can be drawn into a wire two hundred thousand of which, laid side by side, would cover but an inch. To make this finest wire, however, special methods are used, the gold being drawn out in a wire surrounded by platinum, after which the platinum is dissolved off by nitric acid.

Gold has been so greatly valued because of its many useful properties. It readily mixes with other metals, and is made into alloys, giving compounds widely different in their properties and adaptable to many uses. Gold resists the action of nearly all acids, being soluble only in a mixture of nitric acid and hydrochloric, one of the first, and three or four parts of the latter. This mixed acid, owing to its power to dissolve gold, received the name *aqua regia.* Other solvents will dissolve it, but only when the gold is in a very fine state. Its use may be generally stated under a few main classes. Immense quantities are used in coinage; in jewelry, especially in plating on other metals; in the coloring of china and of glass, and of artificial gems; it is used in gilding, as on picture-frames, books, signs, and ornaments, and can readily be deposited from the electric bath. Gold is widely used in photography, though in small quantities, being deposited in a fine state upon photographic prints in toning them. A by no means small quantity is used by dentists, and a few compounds of gold are employed in medicine. Gold, to unknown amounts, is hoarded by individuals and by States, especially in the East, and is kept on deposit in all civilized countries to insure the redemption of other forms of money. Chemically, gold compounds are useful mainly in coloring, or gilding pottery and in the preparation of stained glass.

SILVER

The metal, silver, is white, harder than gold, softer than copper, and especially notable for its brilliancy when polished, though less so than gold. Silver also is ductile and malleable, readily converted into wire or thin leaves. In weight it has a specific gravity of 10.5, being heavier than copper, lighter than lead or gold. It is the best conductor of heat, except gold, of any of the metals, and almost the best conductor of electricity. It has been considered less valuable than gold, simply because it is more plentiful and because, owing to its great affinity for sulphur, it tarnishes readily in the air, taking on a thin coating of the black compound, sulphide of sulphur. Advantage is taken of this to darken silver by the sulphide, when it goes by the name oxidized silver. For use in coins and jewelry, silver is alloyed with copper, coined silver being $^{900}/_{1000}$ pure silver in the United States, and $^{925}/_{1000}$ in England, the latter being "sterling silver."

Silver ores are very widely distributed, and are usually compounds with carbon and sulphur, the commonest being silver sulphide. Vast quantities of silver are produced by the mines, especially in the United States.

Another common compound is the chloride of silver, or horn silver. A very large part of the silver is found in combination with lead, and most lead ores contain it in paying quantities. The most important processes in silver mining are those used to separate lead from silver. One method is to heat the ore that contains the two metals, which causes the lead to crystallize when the heated solution is allowed to cool. Since the lead crystallizes earliest it can be dipped from the solution, leaving silver that contains only a small amount of lead. The solution is next, while heated, exposed to a blast of air and flame. This causes a combination of the lead with oxygen, and the blast carries off the lead oxide thus formed, leaving the silver pure.

By another method zinc is added to the solution of lead and silver, combines with the silver to form an alloy, and then as the solution cools, this alloy, zinc and silver, rises to the top, being lighter than the lead, and may be skimmed off. A small portion of lead still remains, but by being heated, this melts before the alloy and is allowed to run off. A further heating of the zinc and silver alloy vaporizes the zinc and leaves the silver pure.

Of the compounds of silver one of the most important is silver nitrate ($AgNO_3$). It is formed by dissolving silver in nitric acid. It is used for sensitizing photographic paper, for making indelible ink, and for certain dyes. When crystallized and moulded into sticks, it is known as lunar caustic, and used for cauterizing wounds. Although silver nitrate is sensitive to light, and turns dark when exposed to it, especially if accompanied by certain other substances, yet it is usual to convert silver nitrate into silver bromide, or silver chloride, in making photographic plates, since these compounds are even more sensitive. Silver chloride is prepared from silver nitrate by adding chlorine, either as hydrochloric acid, chloride of sodium, or any other soluble chloride. It is believed that light so affects silver chloride, or silver bromide, as to separate the silver into a pure and finely divided state, and this action is greatly hastened and increased by the use of the developer in photography.

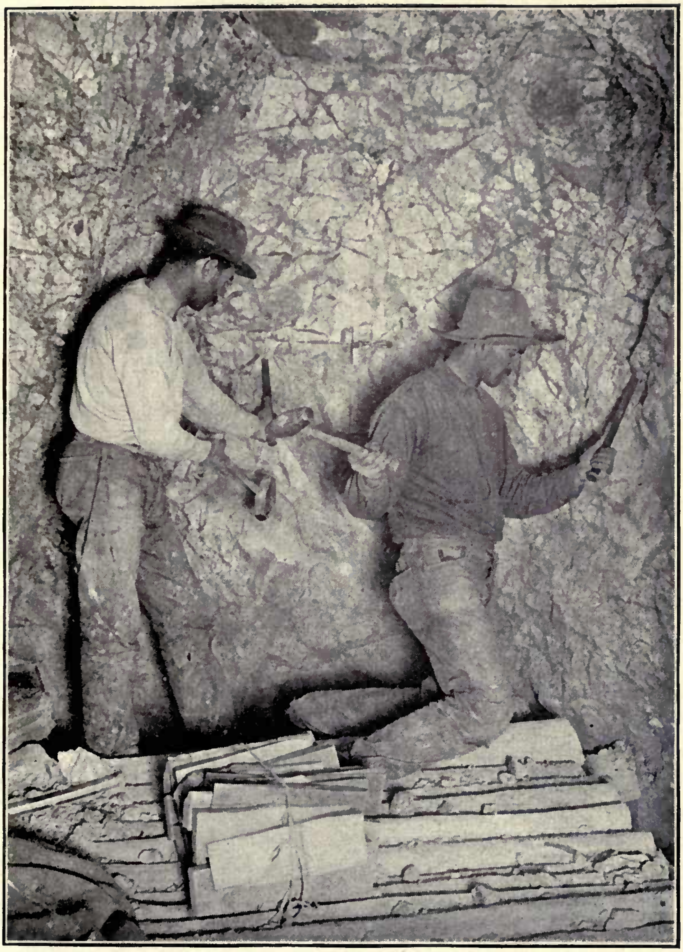
The chemistry of photography has been most exhaustively studied, and the result has been the creation of an endless number of processes for making plates, developing them, and especially in devising methods for printing. The underlying principles, however, are, in the main, the same in all the silver processes. The fact that silver chloride is soluble in a solution of sodium thio-sulphate, or "hypo," is what makes it possible to fix photographs, since by use of the hypo bath the silver compounds that have been converted to metallic silver may be washed out, leaving the metallic silver unaffected.

Silver forms alloys with mercury, platinum, zinc, gold, copper, and these alloys are employed for various purposes in the arts.

COPPER

This metal, owing to the fact that it is found in ores from which it is easily separated by the use of heat alone, was one of the earliest discoveries of mankind, and seems to have been the first of the metals to be useful in the making of tools and weapons. It was plentiful in many parts of the world, greatly valued as an ornament, readily made into utensils and into wire, and so became one of the earliest means of civilization. Besides, copper was very early found to be changed into an even more valuable substance when mixed with a small portion of tin or other metals. Thus it came about that just after the discovery of this earliest metal in its simpler compounds, mankind entered upon what is known as the Copper Age, and, later, the Bronze Age, which periods in development followed upon the earliest Stone Age, that in which man's best tools and weapons were made of polished stone.

Copper is red, and is very flexible, though not in its pure form elastic. Its specific gravity is 8.9, atomic weight, 63.1. Excepting silver and gold, it is the best conductor of electricity. Although it tarnishes slightly in air, a compound is thus formed on the surface which protects the metal from further action. It is readily beaten into sheets or drawn into wire. When copper is heated and plunged into water, it becomes not stiffened, but more malleable. Its name is derived from the Latin, *Cyprium aes,* which meant Cyprium brass. Later, the noun was dropped and the adjective became condensed into cuprum, from which it gets its chemical symbol, Cu. It forms two valuable alloys when mixed with zinc and tin, and these alloys have been of enormous use in the world, that known as bronze, which contains copper, zinc, and tin, long taking the place in the world's history now held by steel. Even after the discovery of steel, this bronze was long used for making cannon, bells, statues, and so on. At the present day the uses of copper are almost innumerable. We can only hint at their general classes. Thus, it is used in constantly increasing quantities in telegraphing and telephoning, both of which require



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DRILLING COPPER ONE MILE UNDERGROUND.

copper wire in their circuits; in electric lighting and motors, and in electric railways. Copper is used also in all forms of printing, for the making of electrotypes and half-tone plates. It is of endless use in household utensils, for sheathing ships, for coinage, nearly all of the cheaper coins containing large parts of copper. In its alloys, the list of which is constantly growing, copper enters into jewelry, all brasswork, all substances composed of gun metal, phosphorus, bronze, aluminium bronze, manganese bronze, all anti-friction metals, and so on. Our nickel coins, so called, are three-fourths copper.

The compounds of copper are usually very poisonous. There are two main classes of these, known respectively as the cuprous and cupric. These two salts of copper are distinguished by the fact that one of them contains one half as much copper as the other, the cuprous salts containing the most.

An important compound of copper is the copper sulphate, $CuSO_4$, called "blue vitriol." This consists of blue crystals, soluble in water, and the solution is much used where a simple compound of copper is necessary in chemistry. It is a most useful solution in electric batteries. Solutions of copper are also used in spraying-mixtures for killing insects. Pigments derived from copper are Scheele's green, Brunswick green, and emerald green. Dangerous compounds of copper may form when copper articles are exposed to damp air. In moist air, copper unites with oxygen to some extent forming copper oxide, which unites with carbon dioxide in the air, forming carbonate of copper, a green compound. Vinegar (acetic acid) will form with copper, verdigris (or copper acetate) which is a poison. Other foods also may act to form poisonous compounds, so copper vessels should not be used to hold food permanently. They should be tinned, nickeled, or otherwise protected. Owing to the poisonous nature of copper compounds, the use of copper utensils in cookery requires that they be kept clean and polished, and that substances prepared in them be removed as soon as possible, to prevent the vegetable juices from acting upon the copper surface.

The cuprous oxide (Cu_2O) is a bright red powder. It imparts a splendid ruby red color to glass, and is used also in giving color to pottery. Thus it will be seen that the chief use of these copper compounds is as coloring matters, or as poisons.

CHAPTER X

METALLIC ELEMENTS, CONTINUED

IRON

AN important group of metals is formed of the three, iron, nickel, and cobalt. The two latter are seldom found separated. Iron is the most familiar, and, in the sense of being the most useful, may be considered the most valuable of all metals. Though known since very early times, there were long ages wherein copper and bronze played, imperfectly, the part that iron was subsequently to fill. When iron was discovered, probably by the accident of using its ores for the building of fireplaces, it began soon to take the place of the softer metals, copper, and its alloys, and the effect upon the history of mankind was enormous. Iron, because of its superior hardness and stiffness, would take a sharper edge, would keep it longer, could be made into more slender tools with finer points, or, in the shape of wire or thin bands, was stronger and more rigid as a means of binding things loosely together or connecting them firmly.

The chief effect of the change to iron was an enormous gain of time in all mechanical work where tools played a part. For example, in clearing pieces of land from trees or bushes by means of iron or steel instruments, the work could be done in a small fraction of the time required when men had no better tools than those made of stone or even of copper. The nations or tribes that first learned to make weapons of iron and steel had an enormous advantage over their neighbors, and often conquered them. When once discovered, also, iron proved to be much more plentiful than copper, more durable in use, and hence was, in effect, cheaper than the softer metal.

When the history of mankind is divided into ages, as the Stone Age, the Copper Age, the Bronze Age, the Iron Age, and the Steel Age, we shall have covered in the list its whole history down to our own times, for we ourselves are living in the last of these. It is to steel, merely a stiffened and more elastic form of iron, that the greatest triumphs of our times are due. Steel beams give us our towering city buildings, and our railroads, depending for their efficiency on steel rails and steel locomotives; our great ocean vessels have steel hulls and engines; our rivers are spanned by steel bridges; and nearly every part of our lives is dependent upon steel, either for tools and weapons, for the instruments and machinery upon which it depends, or for the mechanisms by which these are made. Every improvement in the making of steel from iron vastly increases the power of mankind. Even the rarer metals most recently discovered largely take their value from the fact that they can be mixed into steel to change its qualities. From the steel pen that writes a letter, replacing the split reed or the shaped quill, to the great airship that is promised for the next war, we are dependent upon iron and its compounds for all the arts of war and peace.

Fortunately, iron is very plentiful, but it is seldom found in a pure state, except in meteorites. The compounds of iron, however, are found throughout nature in rocks, in many minerals, in the soil itself, in most natural springs, and even in the coloring matters that make plants green and blood red. The most abundant compounds of iron are hematite, a combination with oxygen; limonite, where iron is combined with oxygen and hydrogen in a complicated way; magnetite, another compound with carbon and oxygen, and siderite, a compound with carbon and oxygen. Pyrites is a compound of iron with sulphur. There are other sulphur compounds with iron also containing copper.

Iron is particularly plentiful in the United States, mainly in the middle of the country and in the Lake Superior region. Chemists and engineers are constantly at work studying how to improve methods of extracting iron from its ores, and as a result, iron is constantly cheapened.

The usual method of extracting iron from its oxide ores is to crush and grind the ores and then to subject them to intense heat, with coke or coal and limestone, in a blast furnace. Under the intense heat of a forced draft, the hot air causes the oxygen and carbon to unite, leaving the melted iron to collect at the bottom of the furnace. The fires in such a furnace may be kept burning for months, new ore being added at the top while the melted iron is every now and then run into sand moulds and allowed to cool in the form of rough slabs, or bars, called pig-iron.

Iron produced in this way is not pure, but contains considerable carbon, phosphorus, and other elements. Owing to its impurities, cast iron is

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LOADING CARS WITH IRON ORE IN A TYPICAL MINE.

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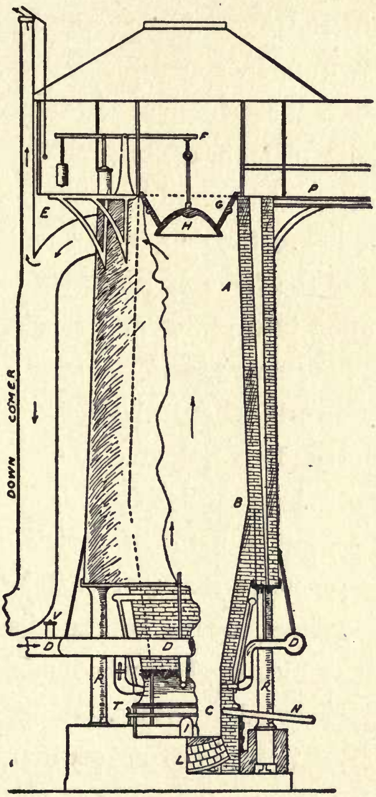
brittle and harder than pure iron would be. If the iron cools rapidly the carbon in it remains combined with the iron, giving it a light color. If



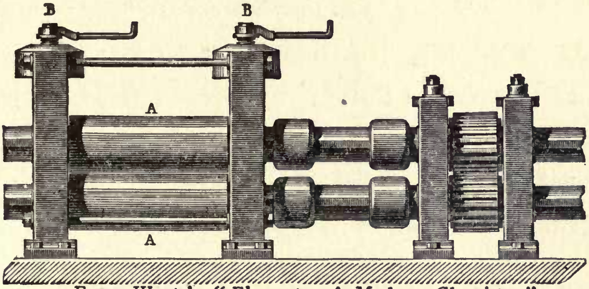
BLAST FURNACE DIAGRAM

Blast Furnace. A, throat; B, bosh; C, crucible where the melted iron collects; D, pipes for hot air blast; E, escape pipe for gases which do not escape through the "down corner"; G, cup; H, cone; N, trough for drawing off slag; T, tuyere; I, hole through which iron is withdrawn.

it cools slowly, much of the carbon remains in the form of graphite, making the iron grey. The latter is the sort ordinarily used in foundries, and is made into stoves, pipes, fences, the heavier parts of machinery, and, in general, all those things in which solidity is more important than elasticity.

A purer form of iron, containing less carbon, is known as wrought iron, and has more of the qualities of steel, being tougher, more malleable, and softer. This is the form of iron seen in iron wire. It may be rolled into plates and sheets, made into rods, chains, horseshoes, and, generally, into those forms of iron-work where elasticity is desirable. It bends readily, does not break as easily as cast iron, and may be readily formed into delicate shapes. It is much used in artistic iron-work. Many things formerly made of cast iron are now constructed of steel, owing to the cheapness and exactness with which steel may be made of qualities to suit the most varied purposes.

Wrought iron is made from ordinary cast iron by heating and thorough mixing so as to burn out much of the impurities. When brought to the right standard, the iron can be rolled and hammered into shapes convenient for use. The process corresponds to that used by the smiths since early times in heating iron in their fires under a blast of air, and then hammering it into shape upon an anvil. Wrought iron contains very little carbon, and as it is the carbon that gives stiffness to iron, it can be readily understood that steel contains a larger proportion of carbon than the wrought iron.



From Wurtz's "Elements of Modern Chemistry."

APPARATUS FOR ROLLING STEEL

To put the matter shortly, a very large proportion of carbon is found in cast iron; a very little carbon is in wrought iron; adding a slightly greater proportion of carbon to wrought iron converts it into steel, which contains more than one and less than the other. The making of steel, however, is greatly modified by the presence of other elements beside carbon, and even by the method of its making.

Steel may be forged, or hammered, into shape, can be welded—which is to say, two pieces can be united simply by being hammered together when heated. It is harder, stiffer, and more elastic than iron, and finds many more uses. Its most valuable and most remarkable quality is its power of being tempered. By heating the steel and plunging it into cold oil or water, it becomes exceedingly brittle and hard. If heated and cooled very slowly, it remains soft and malleable. In order to give it a quality between these two extremes, we have only to heat the steel, cool it until it is brittle, and then to heat it once more. If the second heating be checked at different points, we may obtain all varieties of hardness and elasticity. This reheating is known as tempering. As the steel is reheated it shows different colors upon its surface, and these changing colors tell just what the result of cooling the steel will be at each moment. Experiments have shown that the following colors are produced successively at the following temperatures:

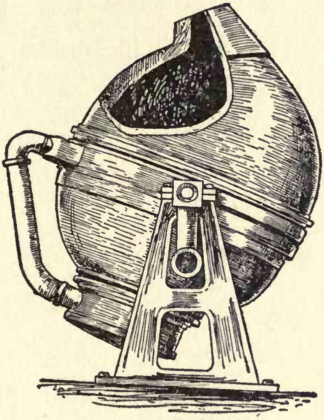
430°	Fahrenheit.	Very pale yellow—Lancets.
450°	"	Light straw color—Razors and surgeons' instruments.
490°	"	Brownish yellow—Scissors and chisels.
510°	"	Purplish brown—Axes and planes.
530°	"	Purple—Table knives.
550°	"	Light blue—Springs and saws.
560°	"	Dark blue—Augers.
600°	"	Blackish blue—Handsaws, hardest drills.

By selecting the proper color, the hardness and toughness of the steel may be governed. For tools meant to be rather tough than hard, the heat is checked while the steel is yellow or brown. To make exceedingly hard tools for drilling or similar uses, the heating is checked only when the steel shows blue.

There are many different substances used for cooling heated steel. The commonest, of course, is water, but many different kinds of oils, metallic solutions, animal fats, and even mercury, are employed to give different qualities to steel. The effects of various elements upon steel may be generally stated as follows: Carbon hardens and toughens the metal; silicon hardens, but lessens the ductility, or the quality of being drawn into wire;

phosphorus makes steel brittle under shock, sulphur makes steel difficult to work; manganese, like phosphorus, seems to increase brittleness, if present in too great quantities; copper seems to do no harm unless sulphur is present with it; aluminium, nickel and chromium toughen steel; arsenic hardens it. It is believed that the best results are obtained when the steel is made entirely of iron and carbon, though small quantities of other elements often render steel suitable for special uses.

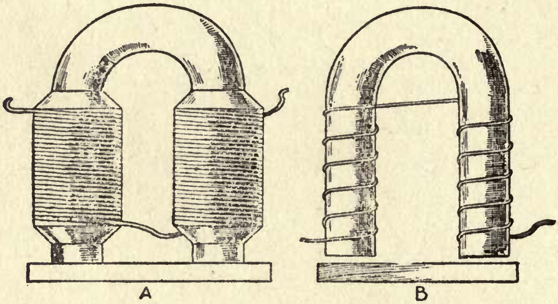
The manufacture of steel, after what has been said, will be seen to consist in securing iron with the right amount of carbon. This can be done either by taking carbon from cast iron or adding it to wrought iron, either directly or by means of adding cast iron to the wrought iron. It is found easiest to add the carbon to wrought iron. This is done either by packing iron with carbon in tight crucibles, or by first burning carbon out from cast iron and then restoring, by adding to cast iron the right amount of carbon. This last is the method used in the Bessemer process, which, invented in 1860, has been most generally adopted and has been found to make steel of any desired quality cheaply and readily. The apparatus in which this is done is known as a converter—a great oval box large enough to hold several tons of melted iron. Through the molten iron air is blown, and the carbon thus burned out,—the oxygen of the air uniting with it forming the inflammable gas, carbon monoxide. This burns at once, keeping the metal melted, and after a few minutes only, the carbon and silicon and other impurities being thus removed, enough cast iron can be added to give any desired mixture of carbon.



CONVERTER

There is another process by which sulphur and phosphorus can be removed in a similar way. A third process consists in burning out impurities on an open hearth by means of a flame containing much oxygen. This is known as the open-hearth process, and is especially useful in making tough and elastic steel, much used where it is meant to stand a great strain, as in cannon, parts of large machines, bridge beams, rails, and building beams.

Pure iron is seven and eight-tenths times as heavy as water. It is a metal having the remarkable quality of becoming instantly magnetic when surrounded by an electric current, and of losing this magnetism as soon as the current is withdrawn. Steel, on the contrary, once magnetized, retains the magnetism. It is believed that these two results depend upon the fact that in steel the particles,



MAGNETS—HENRY'S (A). STURGEON'S (B)

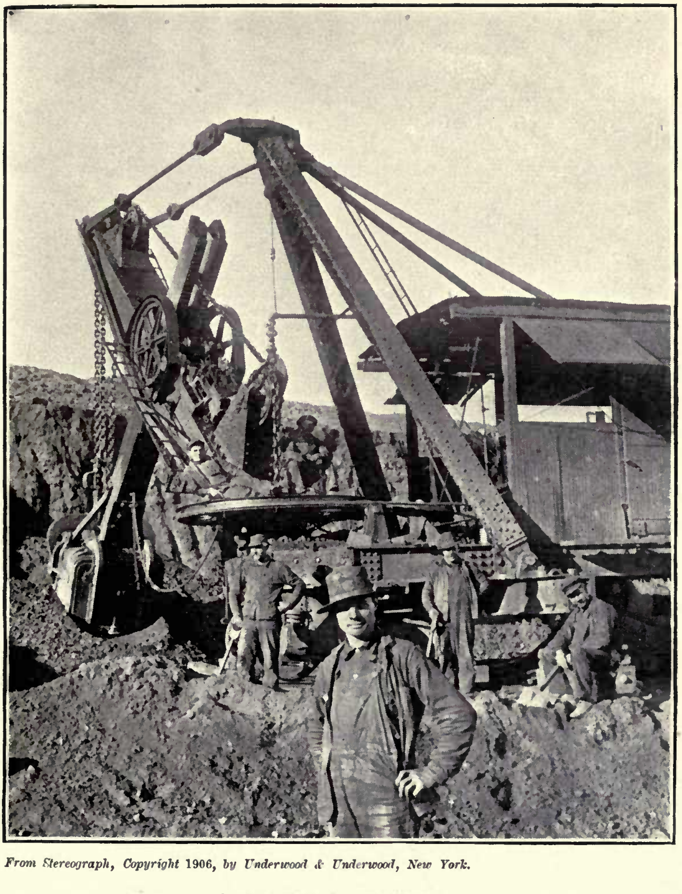
Sturgeon's was the early form, Henry's is the modern improvement, by putting many turns of the wire on spools through which the poles are thrust.

once magnetized, arrange themselves in a fixed position, whereas, in iron, the position of the particles is retained only while the current is passed. The importance of this quality of iron can hardly be over-estimated. Most of the mechanical applications of electricity depend upon it. To it we are indebted for the telephone, telegraph, dynamos, motors, electro-magnets, and all other applications of the electro-magnet.

Pure iron in the presence of water or moisture rusts, that is, forms a ferric hydroxide, which is a complicated compound of iron, oxygen, and hydrogen. Rust formed on the surface of iron does not prevent further rusting, but the rust will proceed until the whole mass of iron is changed to the oxide, because iron-rust and iron, when moist, make up an electric couple, and thus continue the rusting or oxidizing. This will be better understood after reading about electrolysis, later in this book. This makes it necessary to protect iron from the air by means of paints or oils, or by forming chemically upon its surface some compound of iron that will not further combine with the oxygen and hydrogen.

The compounds of iron, like those of copper, are divided into two general classes according to the amount of iron contained in them. They are known as ferrous and ferric salts. Owing to the fact that the ferrous salts readily unite with oxygen, they tend to change in the air into the ferric salts, and this same change can be hastened by the chemist, if he chooses, through the use of acids which bring about a rapid combination with oxygen. The ferrous salts tend to have a greenish color. Among them may be mentioned green vitriol, which is used in the laboratory as a reducing agent, that is, an agent for removing oxygen from compounds. It also finds commercial use in the making of dyes and inks. The inks which are made from iron salts combined with tannin are, in general, of two classes; one which writes black at once, and another which becomes black after drying, owing to the action of the oxygen in the air upon the iron.

Ferric salts tend toward reddish brown and pale violet in color. Thus the ferric oxide is a brownish-red salt which is used to make chalk, and red ochre in painting. It is also a polishing material and gives color to porcelain and glass. These colors vary according to the treatment. The ferric hydrate is yellow ochre, and may be considered sim-



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OPEN PIT IRON MINING.

ply as iron rust. Besides these, there are a ferric sulphate, ferric chloride, and other compounds, from the simplest to those whose composition requires the deepest study. There are hundreds of compounds of iron, but we may say, in general, that besides their uses in the laboratory, they are mainly employed in medicine and in making dyes and inks, and in photography. Their value consists in their coloring property, in their affinity for oxygen, and in the solubility of many of them in water.

COBALT AND NICKEL

Cobalt is a metal discovered in 1733. It has a white, metallic lustre, somewhat resembles iron in hardness, but is more fusible and less magnetic. It has an affinity with oxygen, and thus rusts in the air, though to a less extent than iron. The metal, when pure, is rare, being usually found combined with nickel, with arsenic, or with antimony. It has no especial use yet discovered. It forms salts with nitric, hydrochloric, or sulphuric acid, and these are reddish in color, but become blue when heated, the heat driving off the water contained in their crystals.

Cobalt combined with silica forms a coloring matter known as smalt, a beautiful blue color used in coloring pottery and glass. Cobalt gives other valuable pigments for painters and in the arts generally.

Nickel somewhat resembles both iron and cobalt, but is commoner than the latter. The quality that makes it most valuable is the fact that it retains its brightness in the air, that is, it does not combine with oxygen. Nickel is also used as an alloy with copper, a very small quantity serving to lighten the color of that metal. Alloyed with brass, nickel gives us German silver, a valuable alloy because of its resistance to oxygen, or because it does not rust. Nickel is added to steel, to which it gives the greatest toughness. A little over three per cent. of nickel has been found to make the best alloy of steel for manufacturing armor plates and parts of machinery where extreme strength is requisite. An every-day use of nickel with which we are all familiar is in nickel-plating. This, because of the high polish which nickel takes, not only renders manufactured articles attractive, but it is useful as well, since it protects steel and iron from rusting. Nickel-plating is sometimes applied to zinc, thus giving an alloy which is capable of receiving a very high polish and at the same time is light and durable, while it is easily moulded into articles from sheet metal.

Nickel and iron are likely to occur pure only in meteorites. Like iron, nickel may be welded when free from carbon, silicon, and other impurities. In general properties iron and nickel resemble one another closely, differing only in their affinity for oxygen.

CHAPTER XI

METALLIC ELEMENTS, CONTINUED

PLATINUM

PLATINUM is a very heavy metal, with a strong resemblance in appearance to silver, though it is greyer and does not take so high a polish. It is found in an ore which contains also five exceedingly rare metals in small quantities. These are ruthenium, osmium, iridium, rhodium, and palladium. Iron, gold, and copper are often found associated with these. Most of the platinum ore comes from Russia, though small quantities are found in the United States, South America, and other countries. The name is derived from the Spanish, *plata,* which means silver, and was given because the metal was supposed long ago to be a form of silver ore. The ore occurs in the form of grains or scales, and is at first treated with strong acids to remove gold, silver, and copper. This changes platinum into a solution which contains some iridium. By adding ammonium chloride the platinum is precipitated and may then be hammered into sheets.

Owing to its strong resistance to nearly all acids, platinum is used in the laboratory for making crucibles, stills, dishes, and other vessels to contain acids that would attack other metals. Platinum also resists heat, and may therefore be used in heating compounds when vessels made of other materials would be melted. A very important use of platinum depends upon the fact that when heated, or cooled, it expands and contracts to practically the same degree as glass. Consequently it may be used in connecting electric-light bulbs to electric wires. When the current passes into the bulb it heats the connecting wires, and these expand. If these wires did not expand and contract at the same rate as the substance of the glass bulbs, they would either cause the bulbs to crack, or would admit air, thus destroying the vacuum. But where platinum is used, the joint remains tight. No other metal has been found to take its place for this purpose.

Platinum is used also by dentists, owing to its resisting power to corrosion. For the same reason it is used in jewelry, making a good contrast with gold and furnishing exactly the right setting for certain jewels. Owing to its use in electric work, the demand for platinum largely exceeds the supply, and its price is continually increasing. Platinum is exceedingly heavy, even heavier than gold, being twenty-one times as heavy as water. There are only a few elements which are heavier, such as osmium and iridium.

Platinum conducts electricity, but has a very high resistance, which causes it to give out a large amount of light and heat. Platinum is largely used in photography because of its beautiful black color when minutely divided, and its permanence, owing to the fact that it does not tarnish or oxidize. The chief obstacle to its use for the making of photographic prints is its rarity and very high price.

A peculiar property of platinum, while in the spongy state, or when very finely divided, is the ability to absorb very large volumes of gases. The absorption of gas is so rapid that heat is given out rapidly enough to set fire to the gas. This same property is found in palladium. It is known as occlusion. Owing to this property, small quantities of platinum are sometimes used to make lighters on gas-burners. As soon as the gas is turned on, its absorption by the platinum begins, the platinum is soon red-hot, and the gas takes fire. A pocket cigar-lighter is made on the same principle. A small metal box contains a wick saturated with a liquid that gives off gas readily. Another compartment of the box holds a bit of platinum wire, which, by being thrust into the fumes, becomes heated and ignites the vapor, as already explained.

Of the rarer metals often found with platinum, palladium is sometimes used in the laboratory because of its ability to absorb gases, especially hydrogen; osmium is used in a certain form of electric light and in chemical work under the microscope; while iridium is used in making an alloy to tip gold pens, as it is intensely hard and almost unwearable. Some chemists make a large group of nine metals, which they call the platinum group. In this are included iron, cobalt, ruthenium, rhodium, palladium, osmium, iridium, and platinum. They are grouped together because all are hard, white, not easily melted, and, excepting iron, are not attacked by the air.

To finish the discussion of this group by a few words about the rarer metals, ruthenium is called the rarest of the elements; rhodium also is very rare; palladium has interest for us mainly because of its property of absorbing hydrogen. This absorption increases the size of the mass of metal and takes place, as has been said, rapidly enough to cause a high temperature, as in the case of platinum. Radium is still being studied, and therefore cannot be finally assigned its place in a table or system of elements.

ARSENIC

Arsenic, when pure, is a brittle, grey solid, about five or six times heavier than water. When heated it changes to vapor, with a smell like that of garlic. Its chief practical use is to form poisons for animals, as a drug, in minute quantities in medicine, and to harden lead, a very small portion being added to the lead from which shot is made. Arsenic also occurs in many green paints, enters into the manufacture of glass, and is especially useful in preserving skins from the attacks of insects. Paris-green is a common insect poison, containing a large quantity of arsenic. The symbol for arsenic is As. In spite of the fact that the element arsenic itself is not poisonous, its compounds are among the most poisonous of all substances. Fortunately it can be readily detected, even in the most minute quantities. Arsenic coloring matters are very dangerous. In wall papers or clothing or draperies, the arsenic may become detached and be absorbed or breathed as a gas. For this reason brilliant green colors in food, dyes, wall papers, and so on, should be avoided unless they are known to be free from arsenic. Symptoms of arsenic poisoning are dryness of the mouth, headache, and nausea, or, if the poison takes the form of a gas, muscular pains and inability to sleep.

ANTIMONY

The symbol for this element is Sb, from the Latin name *stibium*. It is a silver-white solid, found usually in crystals, a little heavier than arsenic. As it expands instead of contracting when cooled, antimony is added to type metal so that the metal on cooling may fill up the tiniest parts of the mould, forming perfect type.

Antimony is also used in the alloys known as Britannia metal and Babbitt metal, and with bismuth, it forms a very effective thermopile. The thermopile is a means of causing electricity to be produced by heating or cooling two metals joined together, an arrangement often found useful in certain apparatus. Excepting these uses, antimony is found commercially useful only in making some kinds of matches and fireworks, and in the preparation of rubber. It is also one of the constituents of tartar emetic.

LEAD

Lead has been known since the earliest days, being called *plumbum* by the Romans, which gives us our symbol for the element, Pb. It is occasionally found pure in nature, but the most abundant ore is galena, which is lead sulphide, and it is from this ore that the lead of commerce is derived. A large quantity is produced in the middle west of the United States. The ore is subjected to heat in a furnace, and the metal readily separates from the sulphur, which combines with oxygen, or sometimes with iron, which is added for the purpose. Lead is also separated from the galena by the electric current, using sulphuric acid as the electrolyte.

It is a bluish metal, showing a bright surface when cut or scraped. It soon tarnishes in the air by forming an oxide, but once formed, this coating protects the metal within. It is the heaviest, except mercury, of the common metals, and melts readily at a low heat.

The salts of lead are poisonous, and if taken into the body gradually accumulate there, with serious effects. Commercially, lead finds its greatest use in piping, made by forcing lead through a hole in which is a core. This piping is used both for conveying water and to cover the wires that conduct electricity. It is also used, of course, in the making of shot and bullets, though these are nowadays often alloys.

To make type metal, lead is alloyed with antimony, because the latter metal has the property of expanding while cooling, as already mentioned. Massicot and litharge are the monoxides of lead (PbO), and are used in oils and varnishes and in making glass. Red lead (Pb_3O_4) is used in making glass and as an artists' paint. The cheaper form of red lead is often used as a priming, in painting, to protect iron structures from rust, and also to make tight the joints of lead pipes. The old name was minium, and the use of this red color by illuminators for initials in books written by hand gives us the word "miniature," applied at first to small ornamental pictures that went with the red initial letters.

White lead is a lead carbonate ($PbCO_3$). It is a heavy white powder, used as a white paint when mixed with linseed oil. It is prepared by many processes, the older processes being very slow, but the electric process is rapid and said to be as good.

Sugar of lead is the lead acid formed by the action of acetic acid upon lead, or lead oxide. It is exceedingly poisonous. There are many other compounds of lead, but they are less important.

TIN

Tin is one of the oldest of the metals, the mines in the British Islands having been known long before the Christian Era, being visited by the Phœnicians. The islands, indeed, were known as the Cassiterides, or Tin Islands, from the Greek word Κασσίτερος. The Latin name for the metal is *stannum,* which gives us the symbol Sn. While softer than zinc, it is harder than lead, and a little over seven times as heavy as water. As it is not easily affected by air or moisture or weak acids, a coating of it is used to protect sheet iron, which forms "tinware." The so-called tin foil should be made of hammered tin, but is usually formed partly of

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lead. Tin is used largely to make alloys with copper, and lead with tin forms pewter and solder.

When a bar of pure tin is bent, it produces, owing to the crystals present, a peculiar noise, known as the " cry " of tin.

MERCURY, ZINC, MAGNESIUM, AND CADMIUM

Mercury is the metal that exists in liquid form at ordinary temperatures. Its symbol is Hg, and its common name is, of course, quicksilver. The symbol comes from the Latin name, *hydrargyrum,* which means water silver. Mercury has been known since the earliest days, and occurs mainly in the important ore, cinnabar, especially abundant in Spain. The mercury is extracted from this ore, which is a sulphide, by heating, whereupon the sulphur combines with oxygen, and the mercury is changed to vapor, and may then be condensed. Pure mercury is a bluish white liquid, which boils at 674 degrees Fahrenheit. It does not freeze until the temperature reaches about 35 below zero. It is nearly fourteen times as heavy as water. It is one of the elements each molecule of which consists of a single atom. It is used in the extraction of gold and silver from their ores, to make barometers and thermometers, and especially in the so-called mercury pump, used to extract the air from incandescent bulbs. When made into an alloy, it is called an amalgam. Amalgams are used in tinning, gilding, and silvering, to cover the zinc plates in electric batteries, and to protect metals from rusting. In medicine, mercury is a most important drug, especially used in blood-poisoning and in liver troubles, being the chief constituent of calomel. Mercury is, however, a poison to the system, producing palsy, salivation, and other symptoms. Another use is the silvering of mirrors, for which purpose mercury is made into an amalgam with tin. Amalgams made with silver or platinum are used by dentists for filling teeth. Mercury chloride, the common name for which is corrosive sublimate, is a white solid, with acid properties, exceedingly poisonous, but used in minute quantities in medicine, being one of the very best antiseptics. Mercury iodide is used as a paint.

The sulphide, cinnabar, occurs in two forms, red and black, the latter being artificial. The red, when pure, gives the brilliant "vermilion," whose uses are manifold. Fulminating mercury is a complicated compound, with the formula, $HgN_2C_2O_2$. It is a white powder, used to make percussion caps, and so easily exploded that the very slightest disturbance, such, for example, as removing the cork from a bottle, is sometimes sufficient to set it off.

ZINC

With mercury are often grouped zinc, magnesium, and cadmium. Zinc ores are combined with sulphur, carbon, silicon, and oxygen. The sulphur compound is found in Missouri and Kansas. The zinc is separated from its ores by heating with charcoal, and the liquid zinc is then vaporized and cooled again to separate impurities. It is a shiny white metal that is brittle until it has been worked over. Though it does not tarnish readily, gradually it darkens in the air.

It finds an extensive use in electric batteries, for which purpose it is coupled with copper or with carbon to make the two poles. Sheet zinc is used to line tanks, or to protect walls or floors from heat. The coating of zinc on iron is used to protect it from rust, and such iron is commonly known as galvanized. Zinc enters into many alloys, and also forms the basis for a white paint. It is very permanent because it does not change when exposed to the air. Zinc sulphate is used in medicine, and also in dyeing and cloth-printing. A wide use of zinc in sheets is in roofs, gutters, ornamental cornices, and so on. A chloride of zinc is injected into wood to prevent decay. An oxide of zinc, already spoken of as being used as a pigment, is also sometimes used for tooth filling. The symbol of zinc is Zn, and it melts at a little above the temperature at which coal lights, being less fusible than lead or tin, but far more fusible than iron.

CADMIUM

Cadmium, the symbol of which is Cd, is often found in zinc ores, from which it is separated before the zinc, being more easily made volatile. It resembles tin in color, and is more malleable and ductile than zinc. It melts at a slightly higher temperature than lead. The chief use of cadmium is in cadmium sulphide, a bright yellow solid used as a pigment and making a brilliant and permanent yellow. Cadmium also forms certain alloys that melt readily.

MAGNESIUM

Magnesium (Mg) is found very widely distributed in its compounds, and is so abundant as to form over two per cent. of the earth's crust. In the upper Mississippi Valley a magnesium calcium carbonate, called dolomite, and resembling marble and limestone, forms vast deposits, and even mountain ranges. Magnesium is found in many silicate rocks, and also in sea water and mineral springs. It is named from its occurrence in Magnesia, a district of Thessaly. Its melting point is high, about 806° Fahrenheit. Magnesium occurs in talc and in soapstone, a more compact form of the same mineral, also in asbestos. Magnesium sulphate, known as Epsom salts, is found in many springs in the neighborhood of Epsom, England. Some compounds are used as fertilizers, and in making medicines.

Two striking properties of magnesium are its lightness, and the fact that it burns readily, giving a most brilliant light, often used in flashlights for photography, and also formerly in magic lanterns. Phosphates of magnesium occur in the bones of animals, in grains, and seeds, and in guano, which renders the latter valuable as a fertilizer. Magnesium oxide, often called magnesia, is used in medicine as an antidote for poisoning with mineral acids, and since it hardens on exposure to the air when mixed with water, is a useful compound in forming artificial stones, or gems.

ALUMINIUM OR ALUMINUM

Next to oxygen and silicon, compounds of aluminium are most plentiful in the earth's crust. Feldspars, mica, clay, and slate are the commonest minerals in which aluminium is found. It forms about eight per cent. of the earth's crust, and is therefore the most abundant of the metals. The metal was first obtained, in 1827, by Wöhler. It was prepared also in 1854 by Ste.-Claire Deville. Sir Humphry Davy, on account of the relation of this metal to alum, proposed that it be named alumium. Later it was called aluminum, and finally aluminium. Although such vast quantities were known to exist, the difficulty of extracting it was so great that the metal at first obtained was hopelessly expensive. In 1885, at the International Exposition, a small bar of aluminium was exhibited in a glass case, labelled "The silver from clay." The showing of so large a mass of aluminium was the result of a century's hard work by chemists. The metal occurred in alumina, where it was combined with oxygen. In 1824, Œrsted discovered a way to make a compound of chlorine and aluminium, and from his compound, Wöhler, three years later, by the use of metallic potassium, separated the aluminium, since the potassium combined with the chlorine. The metal as separated was a fine powder, which could not be brought to unite, and this powder remained a chemical curiosity for twenty-seven years.

In 1854, Deville, by a slight change in Wohler's process, obtained a button of aluminium, instead of the powder. But this button cost more than its weight in gold. This same chemist then attempted to decompose the chloride of aluminium by means of the electric current, and thus obtained enough to make a small rod. But in those days the cost of electricity was too great to make the experiment a valuable one.

Deville next used sodium, instead of potassium, and this time obtained a large quantity of aluminium at less expense, part of which was cast into the bar displayed at the Exposition, and another part was made into a rattle and presented to the Prince Imperial. A little more of the same metal was cast into eagles for the flagpoles of the French army, and made into a helmet for the King of Denmark.

No longer ago than 1855, was aluminium thus an almost priceless gift offered to the son of an emperor. To-day, it is cheap enough to meet wide use in kitchen utensils made of precisely the same metal. By 1860 aluminium was being sold for sixteen dollars a pound, but no improvement of importance was made for a quarter of a century. Other investigators tried to improve the process, but it was not until recently that the improvement in electric machines and the cheapening of the current so greatly cheapened the metal.

Without following further the many improvements that cheapened its production, we come to what is known as the Hall process, in which electrolysis made easy and cheap the production of hundreds of pounds a day. The cost of aluminium was quickly reduced to less than fifty cents a pound, and even this was not the limit to the cheapening.

Aluminium is only two and a half times as heavy as water, or about one-third the weight of iron. Its lightness is enormously important, and millions of pounds of the metal are produced a year (nearly all at the electric works at Niagara Falls). It is easily drawn into wire or hammered into plates, though it needs to be annealed frequently. It is an excellent conductor of electricity and heat, and its tensile strength is about equal to that of cast iron. Though readily cast and welded, it can be soldered only after taking great precautions to bring together only clean surfaces of aluminium. It is very slightly affected by the air, and is little affected by many acids. Clay would be a cheap and an inexhaustible source of the metal, except for the presence of impurities, such as iron, silicon, and various metals which make the aluminium brittle.

It is hardly necessary to tell the manifold uses to which the metal has been put, either for ornamental purposes, in making protective paint, or for light metal fixings of every kind. It forms several useful alloys. Corundum and emery are oxides of aluminium. Many of the precious gems are mainly aluminium with coloring matter. The sapphire, the ruby, topaz, amethyst, and emerald (the latter three in the precious "Oriental" variety) are all crystallized aluminium oxide. Turquoise is aluminium phosphate; the emerald is an aluminium silicate; the garnet also contains silicate of aluminium.

Alum is a compound containing aluminium sulphate and potassium sulphate. It is the general type of other compounds resembling it, also known as alums. These are very soluble in water, and find a wide use in dyeing, calico printing, in tanning, in making paper, and in fire-proofing wood and cloth.

An important use of alums is the fixing of dyes in cloth, they being then known as mordants. The mordant is used to wet the cloth before it is printed with the dye, and unites with the dye in such a manner as to form a coloring compound that is permanent, or a fast color.

The clay which will make pottery is aluminium silicate, and ordinarily contains many impurities. According to quality, the pottery is divided into three classes, porcelain, stoneware, and earthenware.

Aluminium fuses at 1292 degrees Fahrenheit. An important alloy is aluminium bronze, which is hard, malleable, and tough as steel.

Aluminium filings, with mercury chloride and potassium cyanide, are put together in such a way as to produce great heat by their combination. Then water is added at such a rate as to reduce the temperature to about seventy degrees. The heat taken up by the water causes it to separate, the oxygen uniting with the compounds and pure hydrogen being set free. This method of procuring hydrogen is in use in Paris for the purpose of preparing hydrogen for balloons.

A modern application of the readiness with which oxygen combines with aluminium is known by the name, Alumino-thermics. A mixture of aluminium and some oxide such as iron rust, when lighted by the magnesium ribbon burns throughout its entire mass with a fierce heat equal to that of the electric arc. This intense heat is used in manufacturing tools of chrome steel, in welding trolley rails, especially the third rails that convey electricity, and also may be used on shipboard to weld large castings such as stern-posts and propeller-shafts, when broken. The thermit can be applied to the break without removing the broken piece from its place, and produces as strong a welding as would be possible in a steel foundry.

MANGANESE

Manganese, although not found pure in nature, exists widely in its compounds, from which, in its oxides especially, it is readily reduced by roasting with charcoal. It has a high melting point, about 3500 degrees Fahrenheit, and is about seven and a half times as heavy as water. When separated, the pure metal is grey, hard, and brittle, with a lustrous surface. Both physically and chemically it resembles iron in many respects, entering into similar combinations. It is especially useful in forming alloys, an instance of which is "spiegeleisen," a valuable form of iron alloyed with manganese.

It forms many oxides which are found in nature. These find employment in dyeing and in medicine. Manganese also in some compounds acts like an acid, and gives rise to manganates and permanganates, of which those of potassium are especially important, being used in the laboratory as powerful absorbents of oxygen, especially the permanganate, which acts with great energy. This latter forms the basis of a disinfectant known by the name, Condy's fluid.

Copper and iron with manganese form manganese bronze, an exceedingly strong and tough alloy, possessing many of the best qualities of steel, valuable because it does not oxidize, or rust. A notable use of this metal is the making of propellers. A compound which is most important and abundant is manganese dioxide, MnO_2. This, when heated, sets free oxygen, and by the use of hydrochloric acid may be made to set free chlorine. This dioxide is much used to whiten ordinary glass, in which it neutralizes the green tint.

CHROMIUM

Chromium also is found in compounds, never free, but traces of it occur in many minerals. Until the intense heat of the electric furnace was used, it was extracted with great difficulty. It is a lustrous grey metal, with a shiny surface when polished. It is slightly lighter than manganese and melts only in the electric furnace. If a large proportion of chromium be added to steel, it makes a hardened steel, useful where the greatest resisting power is needed, as in battleships, safes, and mineral-crushing machines. With potassium, chromium forms compounds, soluble in water, that are valuable as oxidizing agents. Chrome-alum, a compound with sulphur and potassium, is like ordinary alum except that chromium replaces the aluminium. It is used in dyeing and tanning. Lead chromate is the chrome yellow which is the basis of the pigments called chromes. These are yellow, orange, or red.

There are three rare metallic elements that may be classified with chromium. These are molybdenum, tungsten, and uranium. The first enters into certain laboratory compounds. Tungsten is sometimes used as a hardener of steel, and furnishes in sodium tungstate a fire-proofing mixture for cloth. It has recently proved to be exceedingly valuable for making filaments in incandescent electric lights. It needs a peculiar kind of support, and at present the filament is rather brittle, but if care be used it proves to be a most economical and efficient filament, giving better light at a much decreased cost, when compared with the ordinary carbon filament. Over a million tungsten lamps are in use, and about seventy-five thousand a day are being made. It is also used in making incandescent mantles.

URANIUM

Uranium, the heaviest of all the elements, is of great scientific interest, but practically is used chiefly to color glass, to which it gives the color green by transmitted, yellow by reflected, light.

BISMUTH

Bismuth, though a rare metal, is found free in nature and needs only to be purified. Its color is greyish white, with a reddish tinge, and it is brittle. It is nearly ten times heavier than water, and melts at a slightly higher temperature than does tin. Mixed with lead and tin, it forms alloys that melt at very low temperature, and find a use as safety plugs, which at a comparatively low temperature will melt. They are used in steam boilers, to prevent short circuits in electric apparatus, and in devices against fire, as in the automatic sprinklers of large buildings. Bismuth and antimony form an excellent combination for thermopiles, that is, when heated, an electric current is set up between the metals. Bismuth finds a small use in medicine.

STRONTIUM AND BARIUM

Strontium and barium are rare, and resemble calcium in many respects. The pure metals do not occur free, and find no striking uses. Their compounds are abundant. Thus strontium hydroxide is used in the manufacture of beet sugar. Strontium nitrate enters into the making of red fire. Barium salts have a similar use in making green fire. Both of these metallic elements, combined with sulphur, are used in making luminous paints.

GLUCINUM

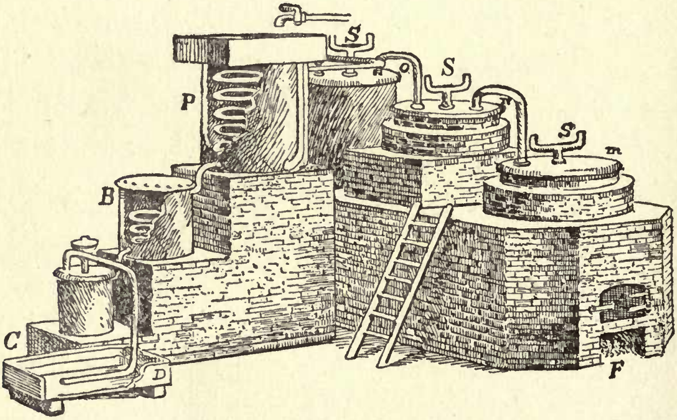
Glucinum is a silver-white metal, light in weight, resembling magnesium. It never occurs free, but is usually found combined with silica, or aluminium, or both. It is found in emeralds, and in the beryl, of which emerald is a species. Its salts have a sweet taste which gives it its name, from the Greek word *glukus,* sweet. Glucinum is also called beryllium.

NITROGEN, PHOSPHORUS, ETC.

We have already, in speaking of the composition of the air, discussed to some extent the properties of nitrogen, especially looking upon it as an inert gas that serves to dilute oxygen and prevent its too energetic action. But nitrogen also plays a most active part in its compounds, and especially in nitric acid and ammonia. It is obtained, as has been said, by removing oxygen from the air, and by taking advantage of the attraction of phosphorus for oxygen, or by allowing liquid air to evaporate—a process which to some extent separates the two gases composing it.

AMMONIA

Speaking now of its compounds, we shall take up first ammonia, which common term is applied both to a gas and to the solution made when the



Preparation of Ammonia

M, *N*, *O*, cylinders containing solutions of impure ammonium chloride, as obtained from coal gas factories, mixed with lime; S, S, S, stirrers to keep the lime from settling; F, furnace to heat the mixture and expel the ammonia; B, B, condensers for cooling the ammonia gas; C, cylinder of pure water to absorb the ammonia and thus form aqua ammonia; D, trough of acid to combine with any fluid escaping from C. In this trough if hydrochloric acid is used, there would form ammonium chloride.

gas is dissolved in water. Ammonia is formed naturally by the decaying of animal or vegetable matters that contain nitrogen. Moisture supplies hydrogen that unites with this nitrogen, forming the gas ammonia. One method of preparing ammonia, formerly much used, was the heating of horny substances, such as hoofs and horns, in a closed vessel, and this gave rise to the name still popularly used for ammonia—spirits of hartshorn. The formula for ammonia gas is NH_3. Since the ammonia gas is lighter than air, it is collected in the same manner as hydrogen—by allowing it to flow upward into a vessel from which it displaces air. Ordinarily, ammonia is prepared by mixing this gas with water, which readily absorbs it. The ammonia in the market is usually that produced by the distilling of coal in making illuminating gas. The solution is correctly known chemically as ammonia hydroxide.

Ammonia is colorless, with a sharp, choking odor that makes it dangerous to inhale. It is only half as heavy as air, and readily separates itself from solution if left exposed. Ammonia can easily be condensed to a liquid form if reduced to a low temperature under compression. When liquefied, it readily evaporates, absorbing much heat as it changes to the gaseous state, and is therefore widely used in refrigeration, as in making artificial ice. The solution is a very strong alkali, so strong as to be called one of the caustic alkalis. It neutralizes acids, causing the formation of salts. These salts have a peculiar chemical formula. In them is found a group of atoms expressed by NH_4. This group acts chemically as if it were a metal, and has been given an especial name, ammonium. It can be separated from compounds, since it at once separates into ammonia gas and hydrogen. This group of one atom of nitrogen and four of hydrogen, is one of those known as a radical, because it is, as it were, a root for the formation of ammonia compounds, seeming to attach other atoms to itself, and yet to retain its formation as a group. This group is found in several salts. With chlorine it makes ammonium chloride; with sulphur, ammonium sulphate. The hydrochloride is used in electric batteries and in several industries. It is commonly called muriate of ammonia, because muriatic, or hydrochloric acid, is used in its preparation. It is a crystalline solid that dissolves readily in water. When heated gently, it turns to vapor, which is condensed upon the cooler surface of the upper part of the apparatus and can then be collected. This method of treating a solution so as to obtain a solid from it is called subliming, or sublimation. The result of so treating ammonium chloride is to produce sal ammoniac.

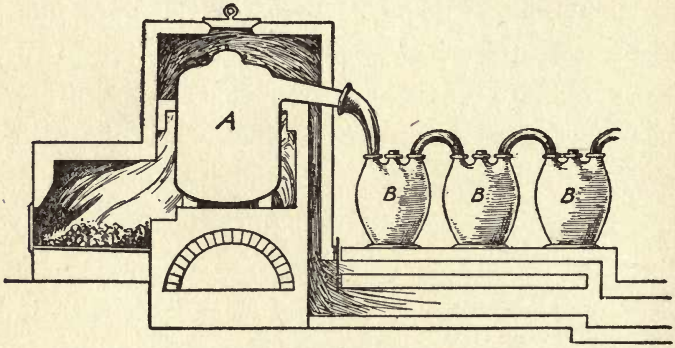
Ammonium sulphate is used in fertilizers; ammonium hydrate is used in the preparation of nitrous oxide, or laughing gas, so called because, when it is used in producing insensibility, during the first stages of its inhalation it causes a certain pleasurable excitement often expressed by laughter.

Ammonia is used as a slight stimulant when inhaled, is very valuable as a cleansing agent, in refrigerating mixtures, and as a dye-stuff.

NITRIC ACID

Nitric acid is prepared by heating sulphuric acid with some salt containing nitrogen; for example, sodium or potassium nitrate. The chemical formula expressing this formation is as follows: $NaNO_3 + H_2SO_4 = HNO_3 + HNaSO_4$, the HNO_3 being the nitric acid.

Nitric acid is very strongly corrosive, causing serious burns to the skin. As it parts readily with its oxygen, especially when warm, and this oxygen combines with most other substances, these are truly burned by the acid. Thus wood or charcoal may be charred by nitric acid as if in a fire. It attacks the softer metals readily, and is therefore used in etching. The etcher covers a plate of copper or zinc with wax, removes portions of the wax where he wishes lines to be bitten into the metal, and then covers the plate with nitric acid, usually diluted. The wax protects parts of the metal, the uncovered lines being bitten in by the eating away



Apparatus for the Manufacture of Nitric Acid

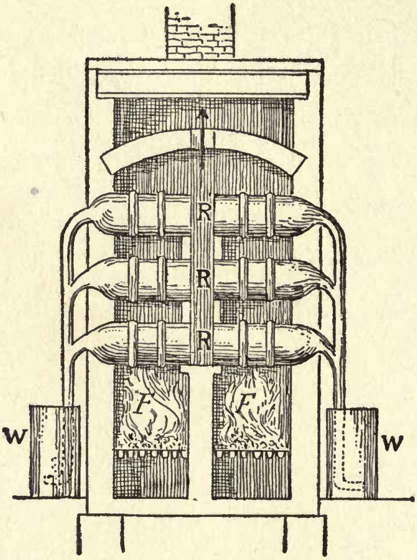
Nitric acid is manufactured on a large scale by heating sodium nitrate and sulphuric acid in a large cast-iron retort A, connected with huge glass or earthen-ware bottles B, B, B, arranged as shown. The last bottle is connected with a tower filled with coke over which water trickles to absorb the vapors which escape from the bottles. The acid vapors are also often absorbed in earthenware or glass tubes.

of the metal by the strong acid. Owing to the activity of nitric acid in combining with metals, older chemists named it aqua fortis, or strong liquid. The action of the acid with the metals forms nitrates, or oxides.

Nitrogen forms five compounds with oxygen, but the only one popularly known is the laughing gas, nitrous oxide, already spoken of.

PHOSPHORUS

Phosphorus, the second number of the nitrogen group, is known to us mainly through the occurrence of phosphorus in nature, and compounds of phosphorus are found in the brain, nerves, and bones of animals. Like sulphur, phosphorus is allotropic, existing in three forms, yellow, red, and black. The yellow phosphorus is not unlike wax, though it becomes brittle when cooled. It readily takes fire in air at a temperature of 34 C., and often at a lower temperature than this will give off white fumes. As is generally known, phosphorus, slightly dampened, glows with a bright yellow light, as may be seen by slightly moistening and rubbing the head of an old-fashioned match in a dark room. Pure phosphorus, owing to its readiness to take fire, has to be kept under water, and must never be handled in the air. The red phosphorus is made by heating the ordinary form in a closed vessel.



Manufacture of Phosphorus

R, R, R, are the retorts into which the mixture of charcoal and phosphorus compounds are put; F is the furnace, and W, W, the water tanks where the vaporized phosphorus is condensed.

The black may be prepared by dissolving red phosphorus in melted lead, and then allowing it to crystallize as it cools.

Phosphorus is used mainly in the manufacture of matches, and also in poisons for destroying animals. While the yellow phosphorus is used for the old-fashioned friction matches, the safety match also uses red phosphorus in the prepared surface upon which the matches are to be struck. Since phosphates are found in plants, especially in the fruits and seeds, vegetation is constantly using up the phosphorus in the soil, and this must be replaced if the land is to be kept fertile, by compounds containing phosphorus.

CHAPTER XII

SOME OTHER METALLIC ELEMENTS

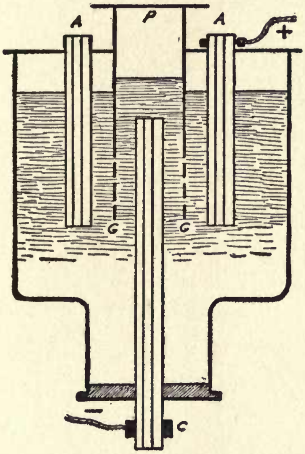
SODIUM and potassium, together with three rare elements, lithium, rubidium, and caesium, are usually classified together, and though they have the character of metals, they seem to differ from other metals in possessing strong alkaline properties.

SODIUM

Sodium, the first of these, for which the symbol is Na, from the Latin name *natrium,* is found most abundantly in common salt, sodium chloride, and in a nitrate, though the carbonate also is common. Something over two per cent. of the earth's crust is composed of sodium, which is about as abundant as aluminium. Sodium, pure, is a silver-white metal, so soft that it is like wax, and can be moulded in the fingers, or readily cut. It is a trifle lighter than water, and melts just below the boiling point of water. At a slightly higher temperature, it burns with a brilliant yellow flame, which is characteristic of all the sodium compounds. To keep it from tarnishing, it is ordinarily immersed in kerosene, or some other liquid free from water—with which sodium will combine by decomposing the water and freeing the hydrogen, while it forms two oxides, Na_2O and Na_2O_2. When put into chlorine, sodium combines with it, forming sodium chloride, or common salt, and by this method Sir Humphry Davy proved the composition of salt.

This compound, sodium chloride, $NaCl$, is found almost universally, being largely produced wherever sea water is evaporated, naturally or artificially. The freezing of sea water excludes the salt from the ice, and the remaining thick solution of salt can be evaporated by heat. Natural deposits of salt, known as "rock salt," are found in many countries, and in these the salt is mined or used impure for curing meat and preserving leather. Water will dissolve about one-third of its own weight of salt, or more, and when the water is allowed to evaporate, the salt crystallizes out in cubes. Common salt often contains chloride of magnesium, which attracts moisture, and in preparing table salt, this is removed.

Sodium hydroxide (NaOH) is used in the manufacturing of soap, in paper pulp, and in many chemical industries. Its common name is caustic soda. Sodium peroxide (Na_2O_2) contains one more atom of oxygen, and is used commercially for bleaching and oxidizing. There is also a sodium oxide. One of the most important com-



Apparatus for the Manufacture of Sodium by the Electrolysis of Sodium Hydronide

The body of the steel cylinder rests within a heated flue. Hence the sodium hydronide is solid in the neck, and serves to protect the joint made by the iron cathode, C, and the crucible. A, A, is the iron anode. A collecting pot, P, dips into the molten caustic soda. As the electrolysis proceeds, the sodium formed at C collects in P, and a wire gauze, G, G, keeps it from mixing with the caustic soda. The sodium is ladled out at intervals from P. The hydrogen, which is liberated, accumulates, also in P, and prevents the sodium from oxidizing.

pounds of sodium is the sodium carbonate (Na_2CO_3), which was formerly made by the burning of seaweed, but is now prepared from the chloride. The carbonate is often called soda, or simply "alkali." It is the compound known as washing-soda. It is used in making soap and in glass manufactures, and in preparing other sodium compounds.

The bicarbonate of soda ($HNaCO_3$), the white powder commonly known as cooking soda, or baking soda, is less soluble in water than the ordinary carbonate. When mixed with cream of tartar, the combination of the two sets free carbon dioxide, and this gas raises bread or cake dough by forcing its way into the paste. The name saleratus is sometimes given to baking soda because of this property, since the word means the aërating salt. Seidlitz powders or Rochelle salts owe their effervescence to the same sort of combination, one of the powders used being soda bicarbonate. In medicine the same compound is often used for indigestion, to counteract acidity.

Sodium sulphate (Na_2SO_4), a white solid, dissolves readily in water, and crystallizes on evaporation into Glaubers salt, used in glass-making, in dyeing, and also, when purified, as a medicine. Sodium nitrate is found in Chili, and is useful as a fertilizer and in the manufacture of nitric acid and potassium nitrate. It is known as Chili saltpeter, and is used in making blasting powder, but makes too explosive a compound for use as gunpowder.

Borax is sodium borate, and occurs in the waters of certain lakes in Asia, from which it is obtained by evaporation. It is also found in California, in solution, and can also be extracted from compounds with calcium and manganese. It possesses valuable antiseptic properties, is used as a food preservative, and, owing to its property of dissolving many of the oxides, is used as a flux in blow-pipe analysis, as is mentioned below in speaking of boron. There are many other compounds of sodium, but these are the most important.

POTASSIUM

Potassium (K) is nearly as abundant as sodium, occurring as a compound in many plants, trees, and animal organisms. It is found as a nitrate in the soil of Ceylon and Bengal. Combined with silica, it occurs in feldspar and mica. Salts of potassium are contained in all soils. Originally, potassium was extracted in the form of potash from ashes obtained by burning wood, leaves, and so on, these ashes being treated with water, and the solution evaporated, leaving what were known as potashes. It is now obtained by electrolysis from potassium hydroxide. It is a soft, white metal, much like sodium, slightly lighter, but ordinarily covered with a greyish coating, owing to the action of oxygen upon it. It melts far below the boiling point of water, and burns with a violet flame, which gives the test for its presence in compounds. Placed upon water, it takes fire, just as sodium does, and burns even more fiercely, giving a violet flame, forming the hydroxide, and releasing hydrogen.

Its compounds closely resemble those of sodium, though it combines even more readily. The pure metal has no commercial uses, is expensive to prepare, and is only a chemical curiosity. An important compound is the potassium hydroxide, or caustic potash (KOH). It is employed in surgery as an active caustic to destroy the skin. It is an alkali of great strength, readily soluble in water, able to neutralize acids, and decomposes many metallic solutions. The nitrate of potassium is saltpetre, or nitre, found in the soil of many warm countries. It is a white, soluble solid, with a salty, cooling taste, giving up oxygen readily, especially to charcoal and sulphur, which property renders it useful in making gunpowder, explosives, and in chemical compounds. Gunpowder often contains as much as seventy-five per cent. of saltpetre, which combines with the oxygen to form expanding gases that give the force to gunpowder explosion.

Potassium chlorate is familiar as a remedy for sore throat. Its value consists in its readiness to part with oxygen, and it is likewise used in the preparation of fireworks. The carbonate of potassium is sometimes called potash, or, when purified, pearlash, and enters into the making of glass, soap, and other potassium compounds.

Bromide of potassium is familiar in medicine, being used as a sedative, and also in photography as a means of slowing the action of developers. Potassium cyanide is often employed in extracting gold, particularly in Africa, and also is used with other cyanides in electroplating. Potassium sulphate is useful as a fertilizer and in preparing alum.

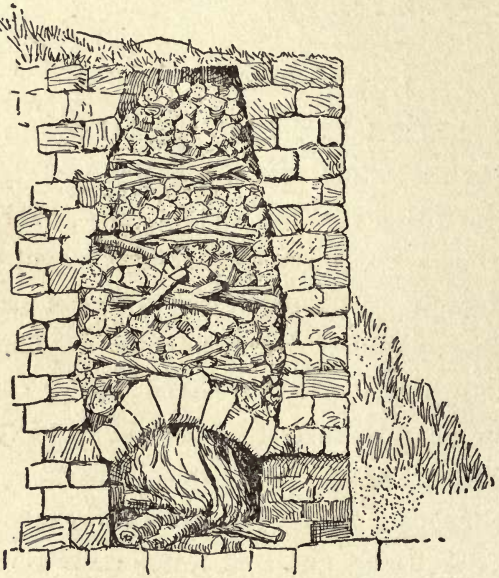
As a rule, the compounds of sodium are cheaper than the similar compounds with potassium, and are therefore apt to replace them commercially wherever that is possible. Now and then the potassium compound will have certain qualities that fit it to a special use, and thus overcome its extra expense as compared with sodium. An advantage in the case of potassium is the fact that its compounds are often more energetic in combination.

LITHIUM, RUBIDIUM AND CAESIUM

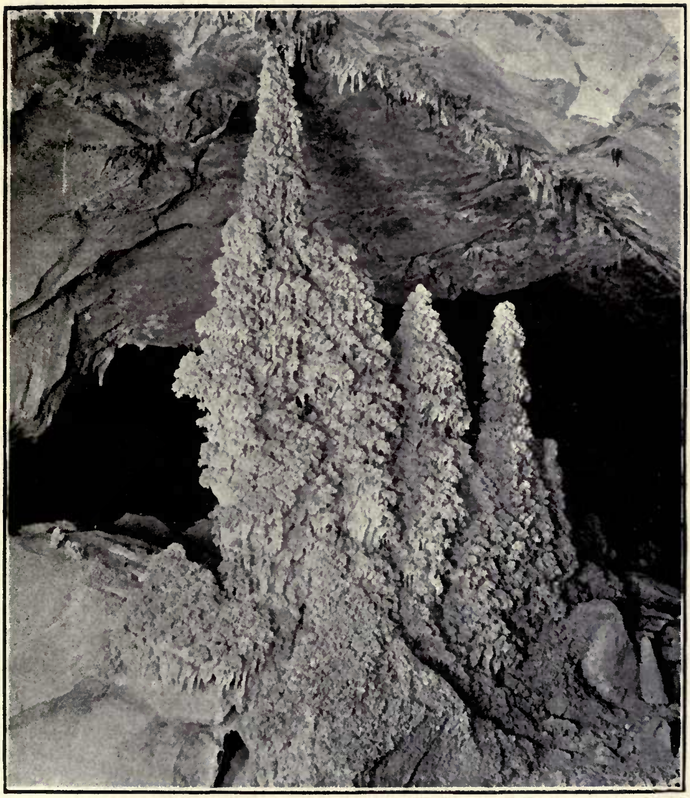
The three remaining metals of this group are not only rare, but are of little use commercially. Lithium is the lightest of all the metal elements. Its compounds are widely distributed, but found in small quantities. Its action is like that of sodium toward water and oxygen, and it unites vigorously with hydrogen, nitrogen, and oxygen. Many of its compounds are almost insoluble, and in this respect lithium is more like magnesium than the alkali metals. The carbonate of lithium and some other compounds are supposed to be beneficial in certain rheumatic disorders. Lithium tends to give flame a bright red color. Rubidium and cæsium are two metals that were discovered by the use of the spectroscope, and received their name from the color of their characteristic lines, rubidium giving red, and cæsium blue lines. They both form compounds much like those of potassium, and have somewhat similar properties.

CALCIUM

The pure metal calcium (Ca) belongs to a class known as "metals of the alkaline earths," being grouped with strontium and barium. While these metals are exceedingly rare, compounds of calcium are exceedingly plentiful, and the compounds of the others are useful and numerous. The metal, calcium, has to be preserved under naphtha, and is only a curiosity. Calcium, however, is found abundantly in marble, limestone, chalk, in shells, bones, coral, etc. The carbonate, or calcide, is the compound that forms these substances, and is found pure in the form of Iceland spar. Calcium sulphate is known as gypsum, alabaster, and selenite. Limestone is found abundantly all over the world, and differs greatly in hardness. Water containing carbon dioxide dissolves limestone, and as such water is very frequent in nature, there are innumerable formations depending upon this property. The water frequently deposits the limestone again, forming the great stalactites and stalagmites. This accounts for the formation of the great Mammoth Cave in Kentucky, and the Luray Cave, in Virginia. From limestone are prepared limes and cements used in building. Forms of rock that owe their formation to shells are very numerous. Examples are the coquina of Florida, and the chalk cliffs of England. Limestone is exceedingly useful in the iron industry,



Lime Kiln



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STALAGMITE FORMATION OF LIMESTONE FOUND IN CAVES.

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being used to extract impurities from the melted iron.

Lime is calcium oxide (CaO). Lime is obtained by burning limestone mixed with charcoal. It is a white solid, soluble in water. It resists heat, and is hardly at all fusible. This makes it useful for crucibles in which even platinum can be melted. Quicklime, when heated by the oxy-hydrogen flame, emits an intense light, formerly called the "Drummond," but now the "calcium light."

The uses of lime are exceedingly important and numerous. Besides forming mortars and cements, it gives us calcium carbide (from which acetylene gas is derived), purifies sugar, makes bleaching powders, is used in dyeing and bleaching of clothing, is used as a disinfectant, and enters into many fertilizers. Limestone with clay forms hydraulic cement, which hardens in contact with water. Such cements as Portland and Rosendale are varieties.

Slaked lime is formed by adding water to lime, making calcium hydroxide $Ca(OH)_2$, and when it is added to water till it forms a clear solution, the solution is called "lime water." When diffused through the water, it forms the solution known as milk of lime, or "whitewash."

Slaked lime is used in purifying coal gas, in tanning, dyeing, and bleaching. Calcium chloride is a white solid, and has the greatest affinity for water. Hence it is used as a drying agent, as it absorbs water from gases or solids. It melts with great readiness, absorbing heat, and is thus a powerful refrigerating agent. A mixture with snow will freeze even mercury.

Calcic sulphate is gypsum, of which we have already spoken, under "Sulphur." Water which holds gypsum in solution is said to be permanently hard, and is particularly troublesome when used in steam-boilers or in cooking or washing. Gypsum is most useful in the preparation of plaster of paris and stucco. It also is useful in forming terra alba, used to give weight to paper. Crude gypsum is a fertilizer.

BORON

Boron has certain likenesses to the hydrogen family, and may therefore be considered with them, though it also differs from them in some respects.

When pure, boron is a greenish-brown powder, which burns when heated in the air, but has no odor or taste. It forms with carbon a carbon boride, which is said to be harder than the diamond.

In nature it is found in the form of boric acid or in its salts. A familiar salt is the sodium salt, known by the common name, borax. The boric acid occurs in large quantities in some localities, being found in jets of steam issuing from the earth. These jets can be condensed in tanks of water, and when the water is evaporated the acid forms crystals. Borax contains a large amount of water, as will be seen from its formula ($Na_2B_4O_7 . 10H_2O$). Borax is often used in the laboratory to dissolve substances for examination. The borax is melted into a bead on the end of a piece of platinum wire. When the bead is colored by the substance to be examined, and then is held in the flame of a blowpipe, the changes of color, and so on, enable the chemist to tell what has been added to the bead.

Borax is a white crystalline solid, used in the manufacture of enamel and glass, and for making artificial precious stones. It is a preservative for canned goods, and is especially good as a cleansing agent in laundries, and is often combined with soap. Having the power of dissolving oxides, borax is placed on metal surfaces meant to be soldered and removes any oxide, leaving the metal clean. It is also used in calico printing, in dyeing, and enters into certain ointments in medicine. Boric acid (H_3BO_3) is used in much the same way. It will be seen that this does not contain the sodium found in borax.

SILICON

Silicon is sometimes classed with carbon, and sometimes with boron, but as carbon compounds are so complicated, they are best studied by themselves, and the description of silicon may well follow that of boron. So abundant are the silicon compounds that they make up perhaps one-fourth of the earth's crust; and silicon is the most important mineral element, as carbon is the most important in organic life. The commonest compounds are silicon dioxide (SiO_2) and silicates. An example of the family is quartz, or sand. The silicates are usually found as rocks. Silicon does not occur free in nature. The dioxide, called also silica, occurs in a great number of varieties that differ in form and color according to the other substances contained in them or to their method of formation. Rock crystal and amethyst are crystalline silica. Other varieties are chalcedony, carnelian, jasper, flint, and so on. Silica combined with water makes the opal. What is known as petrified wood is formed by the washing away of woody fibre and its replacing by silica. Quartz cannot be melted except in an electric furnace or the hydrogen flame, but when fused, it may be drawn out into delicate threads, which are so fine as to be used in scientific instruments, or may be used like a glass.

Silica is very common in plants, and is used in their construction to give them stiffness and to strengthen joints. Thus the coating of bamboo and certain grasses, as well as the quills of feathers and other such substances, are stiffened by silica. Silicates are the result of the action of some acid containing silicon upon various bases. They make up a large part of the earth's crust, and form many common rocks and minerals, among which we may mention mica, feldspar, clay, slate, and talc. Silicates of sodium, or potassium, dissolve in water, and the solutions are known as water glass, being used in yellow soaps, in cements, and in fire-proofing clothing and wood. It will be remembered that in speaking of carborundum we said that it was a silicon carbide, CS.

One of the most important compounds we owe to silicon is glass, which is a mixture of certain silicates, especially of potassium or sodium. There are a great number of mixtures that will form glasses for various purposes, from the peculiar mixtures used in manufacturing photographic lenses to those used in making glass objects such as ordinary bottles. Colored glasses are made by adding substances to the mixture when melted. Thus ruby glass depends for its coloring upon gold or copper. Yellow may be made by adding charcoal, sulphur, or silver; and other substances may be used according to the color desired.

Silicon is, next to oxygen, the most abundant of all the elements, since it is found not only in the rocks but also in all forms of clay. There is a certain likeness in many of the compounds formed by silicon and by carbon, the same number of elements of each combining with a like number of atoms of certain other elements.

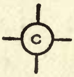
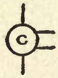
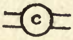
The first glass was made with salts of potassium mixed with the silica; and this glass, known as Bohemian glass, is not so easily melted as that made with sodium. By mixing lead oxide with silica and potash, plate glass is made. This is soft and easily worked, but its high power of refracting makes it useful for lenses generally. Very pure flint glass is the basis from which artificial precious stones are made. Ordinary glass, known as crown, is a silicate of sodium and lime. The cheaper grades of glass contain an iron oxide that gives them a greenish color. Glass has to be carefully and slowly cooled, or annealed, in order that it may not be too brittle.

CHAPTER XIII

CARBON AND ITS STRANGE COMPOUNDS

WHILE it is true that we have already treated briefly carbon and some of its combinations, yet since carbon forms the foundation of all the most important compounds that enter into living organisms, and also of most substances that are derived from them, its chemical story is so wide and varied that it has long been customary to divide all chemistry into two great branches, the second of which is devoted almost exclusively to carbon compounds.

The reason why carbon is thus distinguished from other elements is twofold: first, as we have said, it enters into substances that being in, or derived from, living creatures or organisms, are conveniently considered by themselves; secondly, owing to its nature, carbon forms an almost endless variety of compounds. To explain this second feature we must bear in mind the subject of valency. In discussing the table classifying elements by the periodic system, we shall consider this subject and show how elements vary in regard to it. Carbon ordinarily has the valency 4; that is, it is capable of replacing four atoms of hydrogen (or other monad).* It is also able to remain in combination with any four monads, two dyads, one monad and a triad, and so on. This alone would produce a multiplicity of compounds, but carbon also has the power, like some other elements, of forming combinations in which atoms of carbon are themselves combined, thus making substances in which atoms of carbon are combined more or less with one another. Thus if we suppose two carbon atoms joined, they may hold many more than the equivalent of four monads in combination.

Exactly how atoms are linked together in compounds is not known, but it is usual to represent this linking as if the element had as many links or arms of attachment as it has valency. Represented thus, the carbon atom may be remembered as a circle containing the letter C, and with the four arms of a cross projecting from it, thus:  or  or  and in other similar ways.

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* An element of a valence of *one*, is called a monad; of two valences, a dyad; three, triad; four, a tetrad; five, a pentad; six, a hexad;—and so on, using the Greek numbers to form these words, just as in Geometry the same prefixes are used in the words: triangle, tetragon, pentagon, hexagon, heptagon.

Though we are not to think of this as representing the truth, yet it helps us to follow the linking of carbon, and it is readily seen that if these atoms with their four arms may be joined in a chain, or a ring, we may have compounds with almost any amount of valency, that is, that will admit of putting together elements in endless sets of systems, for to the carbon arms may be attached other elements, each of which has arms of its own, and thus the chain be continued in all four directions.

Nor is all this a mere matter of "chemistry on paper." Carbon atoms do admit of such compounds, and they are known in nearly unlimited number. It is true that some of them may be thrown into classes, owing to the resemblance of the general plan upon which they are built up. But the substitution of a single atom or set of atoms will change a compound into a new substance with new qualities.

The chief elements that enter into all *organisms* are carbon, hydrogen, and oxygen, and in a particular class—the animal substances—nitrogen exists also. The combinations of hydrogen with carbon, alone, are the foundation for an enormous class of substances known as the "hydrocarbons," and these differ greatly, being modified by the amount of oxygen present. When to these three elements we add nitrogen, to get the nitrogenous substances, we have another large class, and by adding other elements to the members of these great classes we vary them into substances that form the material out of which all organized nature is built up.

In animals and plants, all these substances are in a constant state of change. Atoms are exchanged, combinations are made or broken up, substances are destroyed in one form to reappear in several others, and thus the whole great mass of changes which constitute life are carried on. Affecting these changes more or less in organized life are the great agencies, heat, light, and electricity, the action of which must be studied in order to follow the chemical changes among these compounds so universally present.

In the early days of chemistry, it was believed that many of these "organic compounds" owed their formation to some unknown principle acting in living things, and therefore could not be prepared except by nature—that is, not artificially. But in 1828 the chemist Wöhler succeeded in pre-

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FRIEDRICH WÖHLER
B. Germany, 1800. D. 1882.

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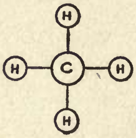
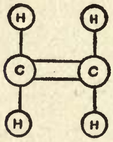
paring the compound urea, till then found only in animals, and never made from inorganic substances. This disproved the belief that all such substances were necessarily beyond the reach of the chemist. Experiments to prepare other organic compounds followed. Many were successful, and as the laws of their making were found out, it came to be believed that these compounds were subject to the same general chemical laws as the others; and that is the belief to-day. It was soon found that the complexity of these substances was due rather to the varied ways in which a few elements were joined, than to the number of elements entering into them.

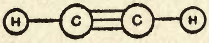
Besides the four elements already mentioned, sulphur and phosphorus were those oftenest found in organic compounds. The cause of the complexity came, as already suggested, from the power of the carbon atoms to unite together; the fact that new atoms may be simply added, or may replace others; that the arrangement of the atoms in carbon compounds may be varied, even when the same atoms are present; and also to the circumstance that in the carbon compounds *radicals* frequently occur—the radicals being atoms combined in groups that often remain together, even when their compounds vary.

It will be seen from this statement that only the most general facts concerning these compounds can be shown in this book. We shall therefore say something about the chief classes of carbon compounds, and add to these merely a mention of some of the commoner substances belonging to this branch of chemistry.

Of the compounds of carbon with oxygen we have spoken already in treating carbon among the elements. The important ones are the gases, carbon monoxide and the carbon dioxide. When carbon is combined with hydrogen, three important gases are produced. These have the formulas, CH_4, known as marsh gas, "natural gas," or methane; C_2H_4, olefiant gas, or ethylene; and C_2H_2, or acetylene gas. By looking at these formulas it will be seen that in the first, carbon is linked to *four* atoms of hydrogen; in the second gas the two carbon atoms are joined to one another by two links, leaving four

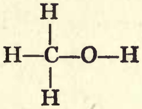






links for the four hydrogen atoms. In acetylene, the carbon is so joined as to link only two atoms of hydrogen.

Remembering what has been said about the power of carbon to make new compounds by change of atoms, we shall not be surprised to learn how these can be varied. To give an example of this—if we imagine one of the H atoms in methane to be replaced by an atom of oxygen, to which the hydrogen atom is again attached, we shall have the compound, methyl alchol.

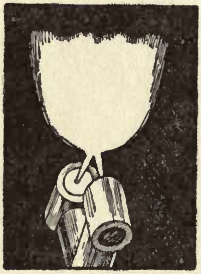


$$\mathrm{H-\underset{\underset{\displaystyle H}{|}}{\overset{\overset{\displaystyle H}{|}}{C}}-O-H}$$

Methane is marsh gas, which rises from marshes, and is found in coal mines under the name "fire damp," and comes from the ground in "natural gas" regions. The second compound (C_2H_4) is known as ethylene, and is found in ordinary illuminating gas, giving the light a yellowish flame. The acetylene formerly formed in distilling wood or coal, is now prepared by combining calcium carbide with water, as in acetylene gas-generators. It gives an extremely brilliant light and an intense heat. All of these, prepared in various ways, are used in illuminating.

ALCOHOLS

We have spoken of methyl alcohol. It is one of a whole class of compounds named "alcohols," of which spirit alcohol is known as ethyl alcohol, and that called wood alcohol is the methyl alcohol. They are made, or derived from, three hydrocarbons called methane, ethane, and propane. The formulæ for these three are CH_4, C_2H_6, and C_3H_8. There is another, C_4H_{10}, called butane. It will be seen on examining these formulæ that in each succeeding case there is one more atom of carbon and two more of hydrogen. Thus these compounds are looked upon as a regular series, and are called the paraffin series. In general, they are compounds in which each carbon atom has been united with hydrogen in all its free links.


Acetylene Flame

The alcohols in general are formed from certain radicals known as ethyl, methyl, and propyl, by replacing a hydrogen atom by the radical hydroxyl (OH). Thus the methane, ethane, and so on, *lacking* the hydrogen atom, give us radicals ethyl, methyl, and so on; and when the missing atom of hydrogen is replaced by OH, or hydroxyl, we have the series of alcohols. Alcohol in Arabic means "the spirit," and these compounds are named from the best known of their class, the ordinary alcohol formed from wine by distillation.

We have already spoken of certain hydroxides, as, NaOH, sodium hydroxide, and KOH, potassium hydroxide. These are metallic hydroxides, and it will be seen that if we suppose the metals to be replaced by the ethyl and methyl radicals, we shall have these transformed into the similar alcohols.

Another class of compounds, known as "aldehydes," are formed when alcohols are oxidized, or, what amounts to the same thing, when the hydrogen is removed. By oxidizing, the hydrogen unites with O to form H_2O. The name, aldehyde, means *al*cohol *dehyd*rated, or deprived of water. From aldehydes are prepared formalin, which is used in dye-stuffs, as a preservative, and a disinfectant. This is a solution of formaldehyde, a gas so called because with oxygen it forms an acid like that secreted by ants and known as formic acid. By the action of chlorine upon alcohol is formed chloral, well known in medicine. This may be decomposed by an alkali, becoming chloroform, the formula for which is $CHCl_3$. If iodine takes the place of chlorine, we have CHI_3, or iodoform, used in surgery as a dressing.

Again we must remind the reader that all these compounds may be greatly varied by changes and substitutions of elements and radicals.

ETHERS

Another important class of compounds are the *ethers,* which also are compounds of C, H, and O. Ether is prepared from ethyl alcohol and sulphuric acid, the result of the operation being to remove from the alcohol the atom of hydrogen in the radical hydroxyl; that is, to change the formula, C_2H_5OH, into $(C_2H_5)2O$. This slight change in chemical composition is the difference between alcohol and ether. The ethers dissolve fats, waxes, and oils, and ordinary ether is used as an anesthetic, as is well known. Here, again, there is a whole group of ethers.

ORGANIC ACIDS

Another group of compounds of the same three elements is known as the *organic acids,* the commonest of which is acetic acid, found in vinegar in a diluted form. Acetic acid may be formed by oxidizing aldehyde, which, as we have seen, is formed by oxidizing alcohol. A set of formulas taken from Newell's "Descriptive Chemistry"* shows very clearly the relation between these three and ethane. Wine is a diluted and flavored alcohol, and is well known to turn into vinegar. This is caused by the oxidizing of the air, which unites with alcohol, forming acetic acid. In nature this turning is caused by a little organism, a formation called "mother of vinegar," which may often be seen, for example, in cider that has stood until it has turned sour.

The action of acetic acid on the bases, gives rise to the salts known as acetates, used in dyeing and making pigments. Paris green is an acetate of copper and zinc.

Among other organic acids are "oxalic," so named from the plant, oxalis; lactic, which forms in milk; malic, tartaric, and citric, in fruits. Many of these acids are in common use in manufactures,

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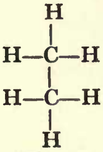
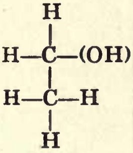


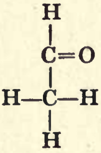
$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H-C-H \\ | \\ H \end{array}$$

Ethane



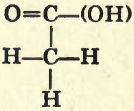
$$\begin{array}{l} \quad H \\ \quad | \\ H-C-(OH) \\ \quad | \\ H-C-H \\ \quad | \\ \quad H \end{array}$$

Ethyl Alcohol



$$\begin{array}{c} H \\ | \\ C=O \\ | \\ H-C-H \\ | \\ H \end{array}$$

Acetic Aldehyde



$$\begin{array}{l} O=C-(OH) \\ \quad | \\ H-C-H \\ \quad | \\ \quad H \end{array}$$

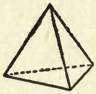
Acetic Acid

and even in the home. Cream of tartar is formed from potassium by the action of tartaric acid, and a tartrate of potassium and antimony is "tartar emetic," used in medicine.

It is, of course, impossible to enter into all the compounds that owe their origin to these organic acids, and, indeed, the slight mention of a few organic compounds here made is designed only to give the reader some idea of the way in which these compounds may be changed by addition or substitution. When it is repeated that the common groups on which such compounds are built may be doubled or tripled, or multiplied, it will be seen that the list of compounds stretches on interminably in a sort of endless chain, and a multitude of these chains exist derivable from methane, ethane, propane, butane, and pentane. And this is only one of the carbon series. Even the larger treatises on chemistry consider only the more important and give the general principles that explain the formulæ of all. Thus we have ethylenes, propylenes, butylenes, amylenes, variations of the alcohols, of the ethers, and so on throughout nature. Besides, we have studies of the compounds, soaps, sugars, fats, anilines—all the derivatives of coal-tar, a chemistry in itself. In addition to these are a whole set of complex substances, such as the albumens, derived from the vegetable or animal kingdom.

It is not to be supposed that any one except a professional chemist will become familiarly acquainted with these higher branches of chemistry. Even these men confine their attention to a few classes of substances and compounds, and refer to books for detailed knowledge of other parts of the science. The general reader must do the same in a larger way. After having become generally acquainted with the principles of chemistry, he will be able, by the use of reference books, to inform himself exactly in regard to matters that may excite his interest or be necessary in his work. Such common substances as butter, milk, olive oil, and soap, are chemically exceedingly complicated. In butter, for example, there are six fats, owing their origin to six different acids. The study of its formation is only possible to the accomplished chemist. Sugar, in chemistry, also covers a whole class of sweet substances, each with its distinct constitution. And when we examine the make-up of such a substance as coal tar, we find that it comprises a complete chemistry in itself—even benzol, one of its derivatives, which is the foundation of many derivatives, gives rise to an endless set of compounds, all characterized by being based upon a ring of the carbon atoms linked together.

Then, too, there is a whole set of theories to explain why compounds of carbon exist in so many forms. A hint of this may be given by saying that the carbon atom is regarded for purposes of analysis as a triangular pyramid capable of attachment to other atoms at each of its four angles, and also of being united with pyramids like itself at various angles and surfaces. Those interested will find the subject discussed under the heading Stereo-chemistry, in various reference books.



Enough has been said to show the reader the endless complexity of this branch of chemistry, and the impossibility of doing more than hinting at its depth.

CHAPTER XIV

Chemical Action and Energy

Thus far, in speaking of the uniting and separation of elements in chemical action, we have thought chiefly of the elements or compounds as they existed before any action takes place, and as they exist when by that action they have been changed. But the most important matter in regard to chemical action, that is, the nature of the action itself, deserves to be especially considered.

The union or disunion of elements is far different from every mere physical mixture, and not only is it different in its results, but especially in its method of action. Perhaps the most striking instance is the commonest—burning. Though we have seen that burning is no more than the uniting of oxygen with other elements, yet we can all appreciate how violent and how remarkable this action is. In the early ages of chemistry the savages saw in the action of fire something so mysterious that fire was worshipped as a god. And even until late in the world's history, fire, on account of its mystery and its usefulness to mankind, was looked upon as holy and almost divine. This will serve to illustrate how striking chemical action may be, and the same is further shown by the whole history of what was known as magic. The properties of phosphorus and of sulphur, the power of acids to dissolve metals, the producing of colored fires, the transformation of solids to liquids and the reverse—all these chemical changes remained for ages the stock in trade of priests, magicians, and wonder-workers.

But only within two centuries men have begun to see that all these mysteries have laws that may be studied and known, and that the knowledge of these laws will enable us to change mysterious happenings to facts as familiar as those of every-day life. Chemical changes a thousand times more wonderful than any known to the "magicians" of old are now practised as part of the daily routine in a hundred great factories. Chemical processes that even the great chemists of the last generation would need a new education to understand are now used in commercial trades and supply us with articles of daily life.

All this has come from a study of the effects produced by chemical action. Those effects are always related to the production or the using up of some form of energy. Usually this energy is evident as heat—either heat given out or heat absorbed. When, in making a combination, chemical elements absorb heat, the result is apt to be unstable, and if the combination be broken up, the same amount of heat taken up by the combining will be liberated. On the other hand, certain combinations give out heat when combining. Combinations so formed are apt to be stable, that is, not easily broken up, since they must, in order to separate, receive from some source the same amount of heat they have parted from in combining.

This will be better understood when it is remembered that we know heat only as a sign of motion, consequently when heat is given out in the making of a chemical combination, it is a sign that the molecules are seeking a position of rest, or ceasing to move to some extent. If heat is absorbed, they are, on the contrary, entering into a state of greater motion, and since they may part again with this heat and return to their state of greater rest, they are naturally in a more unstable condition.

This can be readily illustrated when we think of the explosion of gunpowder. When oxygen is combined with the constituents of gunpowder, much heat is given out by the oxidizing of the elements, and they are thereby formed into new compounds, mostly gases, which have parted with some of their capacity for motion or energy. The study of explosives has shown that they are compounds which are in an unstable condition. Their combination, therefore, produces heat. The addition of an acid to a base is an example of a combination in which heat is ordinarily given out, and a stable compound or set of compounds results.

Sometimes the energy appears in another form than heat. Thus the combination of oxygen with other elements, if rapid enough, will in most cases produce light as well as heat. In the electric battery the action of an acid upon a metal, which is a slower oxidizing, sets free the form of energy known as electricity.

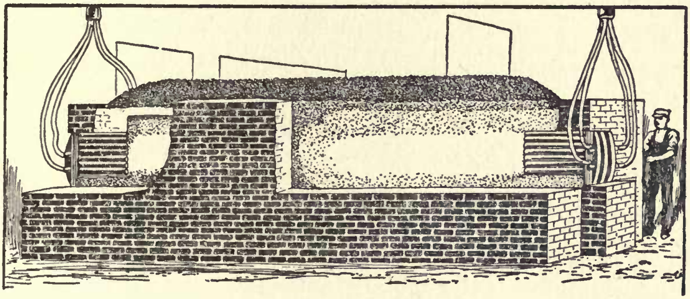
Thus every chemical compound may be looked upon as not only a change in the condition of the substances that enter into it, but also as a releasing or a taking up of some form of energy—either heat, light, or electricity.

The setting free of these is more or less familiar. We have come to expect to see chemical action accompanied by heating, by light, or by electricity, but the contrary—that is, taking up of energy—is less familiar. But only a little reflection will make this opposite action also plain to us.

If light, for example, is allowed to fall upon one of the unstable compounds of silver, such as the bromide, the chloride, or iodide, of silver, as prepared in a photographic film, we know that the effect of that light is to change the composition of the silver compound—to affect it in such a manner that when we apply to the plate a chemical that before would not act upon it, this chemical "develops" the plate, whereas the unexposed plate would not be affected by the developer. This is an instance of the using up of light energy in beginning a chemical combination.

The using of heat energy is more readily recognized. Again we cite the explosion of gunpowder, the burning of a charred stick in oxygen, or the combining of hydrogen and oxygen into water as soon as heat, either a flame or an electric spark, is applied to the mixture. The instances wherein electric energy is absorbed in the making of chemical combinations are numberless. The most familiar of these is electro-plating. In this chemical process the passing of electricity through a solution causes the separation of compounds in solution, and their combination in new forms or in new places.

Of late years many methods of producing extremely low or extremely high temperatures have been discovered, and the proofs that heat plays a most important part in all chemical action have rapidly accumulated. When subjected to temperatures hundreds of degrees below zero, chemicals, which at ordinary temperatures combine rapidly, seem to become inert and have no action upon one another. At extremely high temperatures the opposite is seen—combinations are made of substances that ordinarily can be mixed or brought into close contact without action. To give a modern example, acetylene gas is produced by bringing calcium carbide into contact with water. But this calcium carbide is not produced simply by a mixture of carbon and lime (calcium oxide). The chemical combination of these two is produced solely by the high temperature to which they are raised. At a lower temperature they show no disposition to unite. Another substance due to the high temperature produced by the electric furnace is carborundum. A mixture of coke, sawdust, and sand, on being heated, results in a combination of the silicon and the carbon into silicon carbide, the chemical name for carborundum. This is a substance nearly as hard as the diamond, which is used for whetstones and other sharpening appliances. The fact that this can be chemically made is due to the enormous energy obtained from Niagara Falls and converted by turbine wheels into electricity.



Carborundum Furnace

ELECTROLYSIS

The separation of a chemical compound into its elements by means of a current of electricity dates back into the eighteenth century, for in 1789 two chemists, having sent electric discharges through water in a tube, found that a mixture of gases gradually gathered in the upper part of the tube, and that when accidentally an electric spark was caused in the portion of the tube above the water—one of the wires having become uncovered—these gases disappeared and the water again filled the tube. Study of this experiment led these chemists to believe that hydrogen and oxygen had been separated from the water and then united again into water. In order to obtain even a small quantity of these gases, thousands of electric discharges were necessary.

Later on, when the Leyden jar with its strong charge was invented, from fifty to sixty discharges of electricity through a solution of silver nitrate were enough to cause a visible amount of the metal silver to collect on one of the wires, the negative electrode. Then the current was reversed, making the other wire negative, whereupon the silver deposit gradually disappeared from the first wire and appeared on the other. Meanwhile, Volta had produced his electric battery, and the stronger charges of electricity were used in repeating the experiment of decomposing compounds. With a stronger battery a current sent through river water between brass wires produced bubbles that came in a fine stream from the negative wire. To the astonishment of the experimenters it was soon found out that in the decomposing of water, hydrogen appeared at one wire and oxygen at the other. This was in 1800, and in the same year the experimenter succeeded in decomposing a solution of lead acetate.

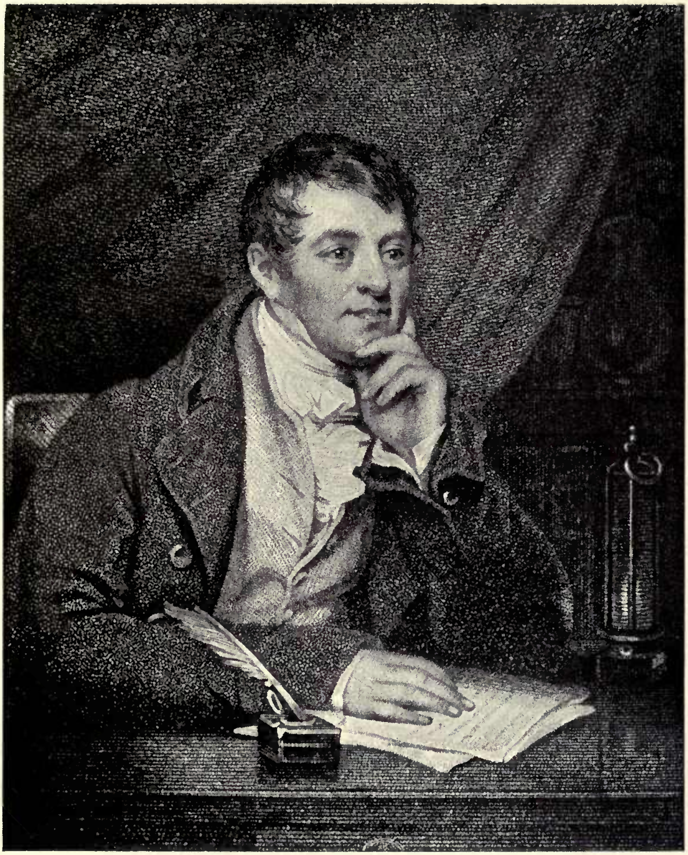
Sir Humphry Davy used two tubes of glass to cover the two separated wires and thereby was able to show that one became a receiver of hydrogen and the other of oxygen, and that these two gases appeared in the same proportion in which they existed in water. There were many attempts made by different philosophers to explain the reason for the separation of the elements at the two poles.

The work of Faraday showed that there was an exact relation between the amount of electricity passing through the solution and the amounts of the separated elements that were set free by the two poles. He gave the names which are still used, and the process he called electrolysis. The liquid decomposed he named the electrolyte; the wire by which it was thought the current entered, he called the anode; the other, the cathode. Those particles which moved about and took new places he called ions, and these he distinguished as anions if they went to the anode, and cations if they went to the cathode. The two poles together have since been named electrodes, and this name is applied to them, whatever their shape or material.

It is now believed that each of the ions is a portion of matter carrying electricity. Because the cation moves with the positive electric current, and the anion against it, the anions are considered negative, the cations positive.

In general, the metals are cations, the non-metals anions; hence metals are deposited at the cathode, non-metals and their compounds at the anode. Hydrogen, however, although considered a non-metal, is like the metals in being electro-positive and appearing at the cathode.

While some electrolytical actions are very simple, others are exceedingly complicated and difficult to understand, since when the compounds or solutions are separated into their elements, various chemical actions take place, producing new compounds. Even the separation of water into its two gases is believed to involve a series of changes depending on the fact that in order to conduct electricity the water contains sulphuric acid, which is decomposed



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SIR HUMPHRY DAVY
B. England, 1778. D. 1829.

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during the electrolysis. It is formed again, however, and thus does not change in amount, serving merely to aid the changes that bring about the release of hydrogen and oxygen.

By his careful measurement of the amounts of elements liberated at the different poles, Faraday established a law which has become of enormous importance in the study of chemistry. He proved "when the same quantity of electricity acts on different electrolytes, the ratio between the quantities of liberated products is the same as between other chemical equivalents." To put the same idea in simpler language, we may think of it in this way: Separating the elements by electricity gives us the same amounts that are found when they are separated by other agents. In other words, it seems that the electricity acts upon the compound only to take it apart, giving us the same results as if we were to take apart the elements by other means.

He also proved that the same result followed always for the *same amount* of electricity, whether it were used strong or weak, or for a longer or shorter time. This enables us to measure precisely the work of the electric current; or by means of the work we may measure the current.

Exactly what takes place in electrolysis is not yet certainly known. It is believed, however, that solutions hold the elements in an already separated condition, that is, the solution contains free atoms or free ions, and these carry the electric current from one pole to the other. When electrically charged, the ions and atoms move, carrying their charges with them to the poles that attract them according to whether their charge is positive or negative. Arriving at the anode or cathode, it is believed they give up their charges, and then act or re-act chemically on one another.

SOLUTIONS

For a long time in the past it was the commonly received belief that when solids were dissolved in liquids, the solid was not changed, but simply became diffused evenly throughout the liquid, though it was early recognized that some substances when so dissolved combined chemically to a certain extent with water. When experiments in electrolysis became common, it was discovered that solutions differed very greatly in the property of conducting electricity. Water, when pure, was so poor a conductor that it might be said not to carry electricity at all. When an acid, base, or salt, was dissolved in water, the solution became, usually, an excellent conductor of electricity, though the power of conducting varied widely. Certain compounds, though they enabled electricity to be conducted, yet could not make solutions that were good conductors.

In order to explain the differences in the conducting power, it was believed that when solutions proved to be good conductors it was because the substances dissolved in them were separated into their component ions. It was found, too, that the conducting power of solutions depended largely upon whether the solution was richly supplied with the substance dissolved, or whether the solution was very diluted. Other points also were noted in experimenting on solutions. The most important one is the effect of dissolving substances in water upon the boiling and the freezing points of the resulting liquids. When compared with the boiling or freezing points of water, it was found that the solution required more heat to raise it to the boiling point, and that its freezing point was lowered. And this raising of the boiling point and lowering of the freezing point is exactly proportioned to the molecular weight of substances dissolved. To illustrate this—if sugar and urea are each dissolved in water in the same proportion, it is found that the raising and lowering of the boiling point and freezing point bear the same relation to one another in the two solutions that the molecular weight of sugar bears to that of urea.

It will be seen that this proportion gives chemists the means of finding the molecular weight of a substance. We may take a substance whose weight is known, and then by experimenting determine the freezing points of the known and unknown substances. We may then make a proportion saying that as the molecular weight of the known substance is to that of the unknown substance, so is the number of degrees of depression of the freezing point of the known substance to the number of degrees of the depression of the unknown substance. This law has been found to give results that agree closely with the molecular weight found by other methods.

But when acids, bases, or salts are dissolved, making solutions that convey electricity, the law does not seem to apply with exactness, and the difference in results seems to indicate that the solutions which conduct electricity are not the same as the solutions which will not conduct it. The conducting solutions, therefore, are believed to consist of more than the mere results of dissolving the substances. They not only dissolve, but dissociate substances, separating them into ions to some extent, as well as into their molecules. As this is a difficult thing to follow, we will put it in another way. All solutions that do not conduct electricity come under the law by which their boiling and freezing points depend upon the molecular weight of the substances dissolved. If the electrolytes, when acids, bases, and salts are dissolved, do not agree with this law, it is because they are different from ordinary solutions. And this difference is believed to consist in these electrolytes being more than mere solutions; that is, there is a separating into ions as well as into molecules.

In Newell's "Descriptive Chemistry" certain proofs of this difference are given. For example, it is there said that *all* chlorides when in solution respond to the same tests chemically. If these differing compounds retain their different character, they would act upon tests differently. The fact that their action, despite the difference in their composition, is the same on all tests, seems to show that when in solution the chlorides have been separated so that the *ions of chlorine* act upon the test rather than the molecules of the chloride compounds.

Another proof of the same thing also given in Newell's book is the fact that this explanation helps to understand the action of all acids, bases, and salts, when in solution. The actions of the various solutions are such as they would be if the compounds had been divided in the solution in the way this theory demands.

A third proof is the easy explanation of neutralization given by this ionic theory. When an acid solution and a base solution are mixed, it is believed that the hydrogen and hydroxyl (HO) unite to form molecules of water. When hydrogen and the hydroxyl ions are thus allowed to combine, the original acid and the original base having lost respectively the hydrogen and the hydroxyl which gave them acid and basic qualities can no longer act as acid or as base, but are neutral.

This theory of solutions is comparatively recent, but it is very helpful in explaining many chemical actions, and there is much evidence to support it in the deeper facts of chemistry. So far as the young reader is concerned, he need remember no more than that when in solution chemical compounds are believed to be in such a state that they consist of particles even smaller than atoms, and that these are able to come together in new combinations. In order to know what these combinations will be, it is necessary for the chemist to know whether the ions set free in each particular compound entering into the solution are positive or negative, that is, whether they tend to go to the positive or the negative pole of wires conducting electricity through a solution. For it is believed that the positive and negative ions, like the differing poles of a magnet, attract one another, and that the ions of the same polarity repel one another, or go to opposite poles.

CHAPTER XV

CHEMICAL LAW. THE PERIODIC SYSTEM

HAVING now made a general acquaintance with the commoner elements and their usual compounds, we are ready to look a little more deeply into the Science of Chemistry, and to become somewhat familiar with its important laws.

In speaking of Dalton in an earlier chapter we told of his theory of molecules and atoms, which led to the law of definite proportions—*elements combine always in fixed proportions*—and the law of multiple proportions, that these proportions if they vary in forming more than one compound of the same elements, are varied in simple proportions.

These laws lead to the finding of numbers showing the combining proportions for each element. And from these, by taking the element oxygen at 16 (or hydrogen at 1) we find the atomic weights for all the rest. Thus for all compounds of any element we have a number which shows that element's "combining weight"; and in its compounds we find either just that weight, or even repetitions of it. For example, we find in oxides either 16 parts of oxygen by weight—or 16 × 2, 16 × 3, and so on.

It is as if each element was in fixed lumps, not to be divided. A whole lump must always be added or taken away in forming compounds. And these imagined "lumps" are atoms.

Consequently a compound must be made up of a certain whole number of atoms of one element, and also a certain whole number of atoms of the other. Therefore the weights of the elements in the compound must compare just as the weights of two groups of atoms compare. Thus HCl is hydrochloric acid. It has an atom of hydrogen, and an atom of chlorine. But the weight of the hydrogen atom is 1.008 and of the chlorine atom is 35.45. So by a simple rule-of-three operation we have the proportion: As 1.008 is to 35.45, so is the weight of the hydrogen to the weight of the chlorine, in any amount of hydrochloric acid. Knowing the weight of hydrochloric acid, we can find both weights.

This is not a lesson book, so we won't do more than show the rule. H_2SO_4 is sulphuric acid, and this contains two atoms of hydrogen, one of sulphur, and four of oxygen. Consequently its parts by weight are:

$$
\begin{array}{lrcl}
\mathrm{H}\ldots\ldots\ldots & 2 \times 1.008 & = & 2.016 \\
\mathrm{S}\ldots\ldots\ldots & 1 \times 32.06 & = & 32.06 \\
\mathrm{O}\ldots\ldots\ldots & 4 \times 16 & = & 64. \\
\cline{4-4}
 & & & 98.076
\end{array}
$$

So, dropping the fractions, there are $^{2}/_{98}$ of H, $^{32}/_{98}$ of S, and $^{64}/_{98}$ of O—from which you can readily put the numbers into percentages or any sort of measures.

From this also you see that each molecule of the sulphuric acid (containing all three) is nearly a hundred times the weight of an atom of hydrogen. Water, being H_2O, its molecule weighs $2.016 + 16 = 18.016$, and is therefore about $^{1}/_{5}$ as heavy as the molecule of sulphuric acid.

Now, since we cannot separate molecules and count them, or get equal groups together, this comparison seems useless.

But remember that matter has several states—solid, liquid, and gaseous. And when we study gases we shall find more help in this comparison of molecules.

Gases are believed to consist of molecules separated and in constant motion—for they may be compressed; gases when unconfined tend to become less dense; and they will mix freely one with another. Their molecules, too, are found alike in all portions. Now in 1660 the English philosopher Boyle showed that it was almost exactly true that a gas always fills the same space under the same heat and pressure, and that gases, at the same temperature, resist pressures alike. Also, heating and cooling affect gases alike in changing volume and pressure, as was shown later by Dalton and others.

The pressure of gases is believed to come from the striking of the moving molecules against the vessel that confines them. Study of the laws that govern the increasing pressure has led to the belief that the loss of volume in compressing gases comes from crowding the molecules nearer together, and that the molecules seem to move in straight lines—though they do *not* repel one another. Heat is believed to increase the velocity of the molecules' motion.

From these facts and theories an Italian professor, Amadeo Avogadro, in 1811 suggested that "all gases at the same pressure, temperature, and volume contain the same number of molecules." This has been generally accepted, and under the name Avogadro's Law has been applied to reasoning out the weights of molecules and atoms, and the explaining of other chemical facts.

Now this law means that at the same temperature and pressure—as in a closed box, kept at a certain heat, and with a gauge that would show when a gas within was at the same tension—the number of molecules is the same, and that they are at the same distance apart. It has proved a generally true law, with only the very smallest exceptions—and these are thought to be explained or at least understandable.

If now we find that equal volumes of gases under equal heat and pressure differ in *weight,* we argue that the molecules differ in weight in that proportion. And thus we can get at the relative weight of the molecules; for if the same volume (at same heat and pressure) is double the weight of another gas, we say this is because the first has the same number of molecules, each of *twice* the weight of the molecules of the second gas.

Only the larger books on chemistry can afford space to explain fully the applying of this law, and the correcting of the slight differences found between the facts found in the laboratory and those the theory requires. We intend only to show the method of reasoning used.

Now, in the case of water, experiment shows that two volumes of hydrogen unite with one of oxygen and form water, making *two* volumes. Supposing we are dealing with 100 molecules in each volume. Then 100 H-molecules + 100 H-molecules join 100 O-molecules, and give us 200 water molecules. But each of these must contain an equal part of oxygen molecules. So the 100 O-molecules have been divided into 200 parts. Therefore each water-molecule contains H (1 molecule or 2 atoms) and O (1 atom). So the expression of the *atoms* in a water molecule becomes H_2O.

But this indicates that each molecule of oxygen is made up of two *atoms*—which may be single particles or groups of electrons.

Though this reasoning is based on the idea of atoms and molecules, yet the relations of the elements thereby found are true, and would remain even if that idea were given up, as is well set forth in A. Smith's "General Inorganic Chemistry."

When we have reduced various compounds to gases at the same pressure and temperature, we can then find out their molecular weight. Analysis shows their composition, and knowing the proportion of each element in the compound, we can find its molecular weight, and its atomic weight—and all these may then be put into tables on the same scale.

The molecular weight of any element or compound means the weight of its molecule compared with that of oxygen or hydrogen. It is found by adding together all the atoms in a compound, each multiplied by its atomic weight.

It is by this method (as well as others) that the weights given the elements in the accepted tables have been found or verified.

Another law, known as that of Dulong and Petit, declares that the atomic weight of any element when multiplied by its specific heat in the *solid* state produces the same number as comes from the same factors of any other element. This means that equal numbers of atoms require the same heat to raise their temperature the same amount. The atomic weight has been defined. The specific heat is the amount of heat necessary to raise the temperature of any element one degree, compared with the amount of heat necessary to raise water one degree.

This law may be put in another way, thus: If the atomic weight of any element be multiplied by its specific heat, the result will be a fixed number. This number varies between 6 and 7 for most elements, and this rule helps to verify atomic weights found in other ways. To give a few examples:

	At. wt.		Sp. heat		
Lithium	7.03	×	.94	=	6.60
Uranium	238.5	×	.0276	=	6.582
Calcium	40.1	×	.17	=	6.81
Gold	197.2	×	.032	=	6.31

If the comparison had been made with another standard than water (for the unit of specific heat) the figure 6. + would be different. But the important matter is the fact that the relation between specific heat and atomic weight is a fixed one with elements when in solid state.

In A. Smith's Chemistry this law is explained thus: In equal weights of elements, there are *fewer* chemical units as the weight of these is greater, and so *few* chemical units require *less* heat to raise them equally in temperature.

All these views of substances, and these laws explaining the making of compounds, the actions of gases, the forming of a certain volume of gas, and so on, come from the general idea that matter consists of elements, elements of molecules, and these of atoms. Whether this be true or not, it has helped us to understand results.

Now we *know* that elements combine in certain weights or "equivalents." We *believe* these are formed by atoms or groups of atoms. Believing in atoms, we can see reasons for the changes in the volumes of gases, for the law of specific heat, and for the differences in various molecules. It is only when we come to study electrolysis and solutions that we need to consider the breaking up of atoms into smaller portions—as has already been suggested.

Accepting, then, the belief in atoms, and their weights, we may examine into the relations of the elements as each made up of like atoms, differing in weight and material from other atoms. If we put the elements in regular order according to their atomic weights, we shall find that there seems to be a certain relation between elements that fall into similar places in the table. The law shown by a table of elements so arranged is known as the Periodic Law.

This arranging of elements was first suggested in 1863 by Newlands, but has been improved by others, notably by Meyer and Mendeléef. The table of elements used to-day is known as Mendeléef's table, and is shown on page 255.

Many other forms may be used in presenting this table, but the general idea is to arrange the elements so as to show that there exist certain groups which when brought together without changing the general order (according to atomic weights) will make it plain that there are many relations between the elements.

Along the top of the table are Roman numbers. These show what is known as the valency of the elements. Valency is defined as the property that elements possess, when entering into compounds, of replacing one another. Thus all elements that have the valency (or valence) "one" may combine with or replace others of the same valency, atom for atom. Those that have a valency of "two" can combine with or replace each other atom for atom, or can replace two atoms of the elements of valency "one," and so on. Thus oxygen has the

Valence O	I	II	III	IV	III V	II VI	I to VII	VIII
He 4 Ne 20	Li 7 Na 23	Gl 9 Mg 24.4	B 11 Al 27	C 12 Si 28.4	N 14 P 31	O 16 S 32	F 19 Cl 35.5	....
A 39.9 x	K 39 –Cu 63.6	Ca 40 –Zn 65.4 x	Sc 44 –Ga 70 x	Ti 48 –Ge 72.5 x	V 51 –As 75 x	Cr 52 –Se 79.2 x	Mn 55 –Br 80 x	Fe 56 Co 59 Ne 58.7 ....
Kr 87.8	Rb 85.4 –Ag 108 x	Sr 87.6 –Cd 112.4 x	Y 89 In 115	Zr 90.6 –Sn 119 x	Cb 94 –Sb 120	Mo 96 –Te 127.6	.... 1.. 127	Ru 101.7 Pd 106.5 Rh–103 ....
X 128	Cs 133	Ba 137.4	La 139	Ce { 140 etc. { 178 Pb 207	Ta 183	(Tungsten) W 184	....	.... Pf 195 Os 191 .... Ir 193. ....
....	....	Ra 225	....	Th 232.5	....	U 238.5	....	....

valence II, and combines with two atoms of hydrogen for each atom of its own. Carbon has the valency IV (or II), and thus combines with other elements accordingly. Many elements have more than one valency, as the table shows.

Elements are named thus, according to valency, from the Greek numbers:

I, Monad; II, Dyad; III, Triad; IV, Tetrad; V, Pentad; VI, Hexad; VII, Heptad; VIII, Octad. The adjectives are, monovalent, divalent, and so on to the end.

In its first column the valence O appears. This means that the elements in that column—Helium, Neon, Argon, Krypton, Xenon, rare gases in the air and elsewhere—show no compounds.

We may know the valence of any element in a compound by seeing how many of its atoms are found combined with those whose valence is known. Thus, in OsO_4 (osmic acid), since O has the valence II, and there are 4 O-atoms to 1 of osmium, we see that osmium must have the valence VIII, as the table gives it.

The table shows how valence changes with atomic weight—rises and falls, periodically. Thus, take the horizontal row beginning with Ne (neon) at

O valence, we have Na (sodium) I, Mg (magnesium) II, aluminium III, silicon IV, phosphorus V (sometimes) or lower, sulphur VI (or lower), chlorine VII (or at times I). It is as if valence grew with weight, till it passed a certain point, and then increased or diminished according to the circumstances. This difference in valence corresponds with the different "equivalents" by which elements form compounds. Thus formula NH_4Cl (ammonium chloride) shows nitrogen as a pentad, combining with an atom of monad chlorine and four atoms of monad hydrogen. Other formulas may be read in the same way.

This relation in valency is only one of the resemblances brought out by this "Periodic System." In general, the properties of the elements increase or decrease with the atomic weights. To perceive this a wide knowledge of chemistry would be necessary, but we may rest satisfied with the general notion that the table brings together the elements in like groups, and that the groups change regularly as we move up and down the columns, or across the rows, to and fro.

Thus in the first column we find rare, inactive elements. In the second, beginning with lithium, the lightest metal, we get heavier metals, with stronger metallic qualities, until we reach copper, silver, and gold. The third column also is increasingly metallic till we reach mercury, and below it (probably) comes radium—which can be rightly put with mercury, the liquid metal, for it seems almost a gas-metal. Column fourth begins with carbon and silicon—the two great foundations of organic and inorganic compounds—and also leads to dense metals—lead and thorium.

Horizontally, from lithium, lightest metal, we proceed to the gas fluorine; from sodium to chlorine; from potassium to manganese—in each case progressing regularly in power to combine with oxygen, or valence toward that element. In each row, the last four regularly lose valence toward hydrogen (if they combine at all). Specific gravities also will be found to be according to certain orders in the table.

But—for its many striking features, chemical treatises must be consulted. We will only pause to say that since the table was first made, it has led to the discovery of at least three new elements to fill spaces that were blanks (scandium, 1879, gallium in 1875, and germanium, 1888). These Mendeléef predicted—with others not yet found. The table also helps to fix doubtful atomic weights, and suggests lines of work to chemists, while it helps to classify the elements, and to confirm discoveries and researches.

CHAPTER XVI

The Story of Chemistry

In very truth the story of chemistry might begin with man's use of fire, or with the earliest use of metals, and we have already seen that speculation about what substances were made of began far back even of the alchemists. But in a true sense, if we wish to know about the modern science of chemistry we can hardly go further back than the English philosopher, Robert Boyle.

Before his century, even, hundreds of compounds were made in the East and in Europe, and there are names of great experimenters such as Albertus Magnus, Roger Bacon, Raymond Lulli, and Paracelsus, who were wise for their time, and held views more or less correct about chemical questions. But with Robert Boyle we first have the idea of widely applicable laws and theories, as against the fanciful superstitions of the alchemists. Boyle was against all such wild notions, and wrote a book pointing out the absurd claims and ridiculously stilted language of the chemists of his time.

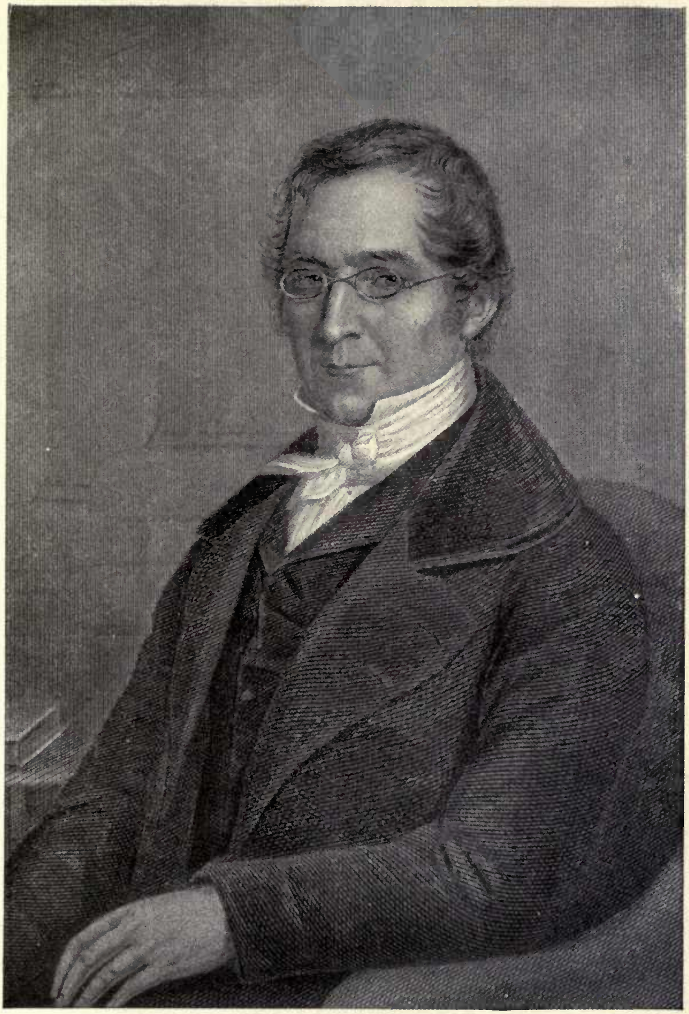
Boyle belonged to a society known as the "Invisible College," a body of men interested in science and the arts, that afterward gave rise to the Royal Society. Boyle championed the view that all matter was made up of smaller particles or bodies, and urged the need of making true experiments, and of setting forth results in plain language all men might understand. With him true analysis of substances began. To him we owe that great principle, "Boyle's Law," discovered by him in 1660—at an unchanged temperature any gas occupies less volume as the pressure is greater—that is, double the pressure and the volume is halved; or halve the pressure and the volume is doubled. This has become the very foundation stone of modern ideas about how matter is made up, for on it is based the theory of atoms, and all that has followed.

Next to Boyle should be mentioned Priestley, concerning whom it is interesting to know that it was Benjamin Franklin's gift of some books on electricity that formed the turning point in Priestley's life, leading him toward science. Though Priestley believed some of the mistaken notions of his day, yet to him, with the great Lavoisier, is due the honor of discovering the composition of air and the element oxygen. Priestley, by means of a burning-glass lens heated "the calx of mercury" (mercury oxide) and found that "air" was given off under a bell-glass. This "air" when tested was seen to be something unknown, and its study was the discovery of oxygen. Priestley also invented the "pneumatic trough" for collecting gases. He separated ammonia into its gases—hydrogen and nitrogen, and made studies upon the nature of combustion—which in his day was believed to consist in the driving out of an imagined substance called "phlogiston." Priestley himself shared this theory. When it was shown that a burned substance sometimes gained in weight (by absorbing oxygen, of course), the believers in phlogiston were forced to argue that it might be a substance of "negative gravity." Oxygen was first called "dephlogisticated air" and nitrogen "phlogisticated air" from this theory.

The next English chemists of note were Black, an opponent of phlogiston, who discovered "carbonic acid gas" (CO_2) and proved its relations to certain carbonates of alkalies and alkaline earths; and Henry Cavendish, the strange, interesting recluse who showed the composition of air, that of water, and who studied deeply into the nature of heat. To him is due the discovery of hydrogen and its uniting with oxygen to form water when burned; and much study of the weight of gases. He showed how the calcium carbonate was dissolved by water containing carbon dioxide, and analyzed nitric oxide. Many amusing anecdotes are told of his eccentricities and his shyness.

Then the great names become frequent, for the true path to chemical knowledge was becoming known. Of Dalton we have told already, and we will only remind the reader of his atomic theory and of the law of multiple proportions, which certainly were first clearly put forth under his name.

The law of definite proportions was at this time much disputed. The great champions were Berthollet and Proust—the former urging that elements combined in many unfixed amounts, and that solutions in all degrees might be regarded as compounds. It was early in the nineteenth century that Proust's party succeeded in showing that mixtures were not compounds, and Berthollet's party were compelled to admit the law. During this con-



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JOSEPH LOUIS GAY-LUSSAC
B. France, 1778. D. Paris, 1850.

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troversy much information was acquired by the disputants, and Dr. Wollaston, famous in electric science, pointed out that elements combined in *multiple* proportions—which completed the rule, "elements combine in definite and multiple proportions"—as we have seen.

The same period saw from the French chemist, Gay-Lussac, a publication pointing out the laws of the combination of gases—that these combine in definite proportions if under the same heat and pressure, and form fixed volumes—which led to the theory announced by Avogadro concerning the reason for the result. The idea that each gas contained compound atoms (molecules) in equal number proved a most fruitful one; and the word "molecule" took its place in the science—but only in after years was the discovery accepted, together with the atomic theory in completer form.

The proving of the theory is credited to Berzelius, a Swedish chemist, and Thomson, a Scotch professor. Berzelius was a great analyst, and tested the theory in his laboratory, making experiments to confirm the combining of elements in fixed proportions. To him is due the system of letters and numbers that enable us to write chemical formulas. This way of writing the results of experiments made the theory evident and widely known.

Next came Dulong and Petit's experiments, showing that the elements had "specific heat" according to their atomic weights, and the discovery by a German chemist that "elements having the same number of atoms to the molecule are disposed to form the same angles of crystallization." This is known as "isomorphism," from *isos,* like, and *morphos,* form.

Dr. Henry Smith Williams in his article, "The Century's Progress in Chemistry" (*Harper's Magazine*), sets forth these facts much more fully, and shows how the atomic theory gradually became an indisputable part of chemistry and the forerunner of the modern science. I have not met with a better brief history of the science and have been greatly aided by his article in this chapter.

With Sir Humphry Davy the union of the sciences of chemistry and electricity began, and almost at once enabled him to make brilliant discoveries. He first separated potassium and sodium as metals from their compounds, and soon in the same way discovered calcium, barium, and strontium. On decomposing water, various substances besides hydrogen and oxygen were set free at the poles, but Davy was able to prove that these were impurities derived from the apparatus.

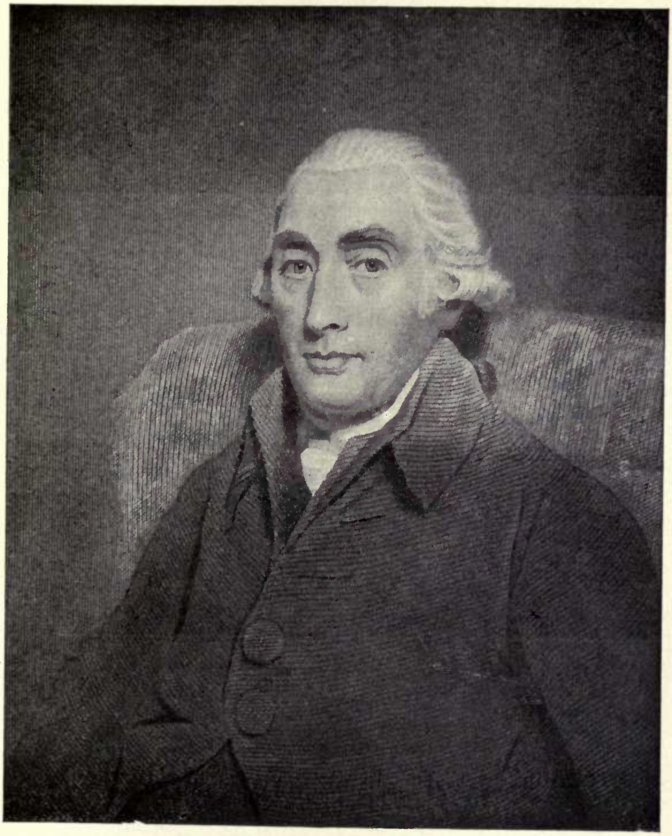
Meanwhile a theory had grown up that all compounds were built up of "binary" combinations of "positive" and "negative" atoms; and this was championed by Berzelius with much ability. It had many facts in its favor, such, for example, as the persistence of the "radicals," of which Gay-Lussac in 1815 had discovered cyanogen (CN) and Ampère had, a year later, discovered ammonium (NH); and for years it was widely accepted.

But later investigations, and increase of skill in the putting together of compounds were to show there was no such general law. Its overthrow came soon after Wöhler's achievement in making the organic compound, urea. This opened the way to the conquest of organic chemistry, and brought about the study of carbon and its compounds that led Liebig to his triumphs in agricultural chemistry—the study of the enrichment of soils and the improvement of food-stuffs; and also saw the entry upon the same field of the French chemists, Dumas and Pasteur.

The studies in organic chemistry revealed many radicals—and many of them were binary compounds. But Dumas was able to substitute oxygen, a "positive" element, for the "negative" element chlorine, and thus to show that the binary theory of Berzelius was not a law.

This led to further investigation and speculation in regard to the make-up of molecules, and brought about the proof that certain molecules must contain more than one atom of the same element. As previously shown, if hydrogen unites with oxygen, making one volume of oxygen + two volumes of hydrogen = two volumes of water-vapor, then the molecules of oxygen must have split into two atoms. If so, the oxygen molecule is O_2.

Similar reasoning with other elements and compounds brought out the theory of "valency," and of the exchange of atoms in the forming of compounds. This was a great gain for several reasons. It explained that there was a limit to the combinations of certain elements with one another. Knowing that the valencies of H, O, N, and C were I, II, III, IV, respectively, it was evident that H and O, for example, could only combine thus, H—O—, H—O—H, H—O—O—H, and that all



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JOSEPH BLACK
B. Bordeaux, 1728. D. Edinburgh, 1799.

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these had used their valencies except the first H—O—, which was less stable than the others, or could not stand alone. It also showed what was meant by the "isomerism" discovered by Liebig and Wöhler—the existing of differing compounds that yet were made up of the same *number* and the same *kind* of atoms. And this led to the close study of how atoms were grouped—as shown by graphic formulas—such as we saw in the chapter on carbon compounds, and also of their action in entering or leaving these groups. Out of these studies have come the ideas of "disassociation"—the tendency of atoms to leave one combination and enter another, and the "reversibility" of chemical combination—or the tendency in many cases of a chemical action to swing back again to its first condition if the elements, molecules, or atoms are not separated or set free in some way at some stage of their movements.

All these views tend to increase the complexity of chemical study, and at the same time to make actions and reactions better understood, and more capable of control. The subject became deeply specialized, and the mass of facts accumulated was so enormous that chemists longed for some simplifying—for general laws that would make the knowledge systematic.

The earliest steps toward this desired end were in an attempt to show that there were not so many elements—that possibly some of them were compounds, and that study might prove there were but few real "elements." It was even suggested by Prout, in 1815, that, since many of the atomic weights were multiples of the atomic weight of hydrogen, it might be all elements were but forms or compounds of this single one. In 1840, Dumas made some attempts toward a possible proof of this theory, but it would not lend itself to proof.

Help came in another way. In 1864 the chemist, John Newlands, showed that if the elements were put in the order of their atomic weights, there seemed to be a law of "octaves," and this led to the studies of the American, Hinrichs, the German, Meyer, and the Russian, Mendeléef—with the result already explained in discussing the Periodic System.

Another help was found in the spectroscope. By the aid of this instrument, and of photography, it was possible to examine and compare the spectra of bodies, not only on earth, but in the heavens.

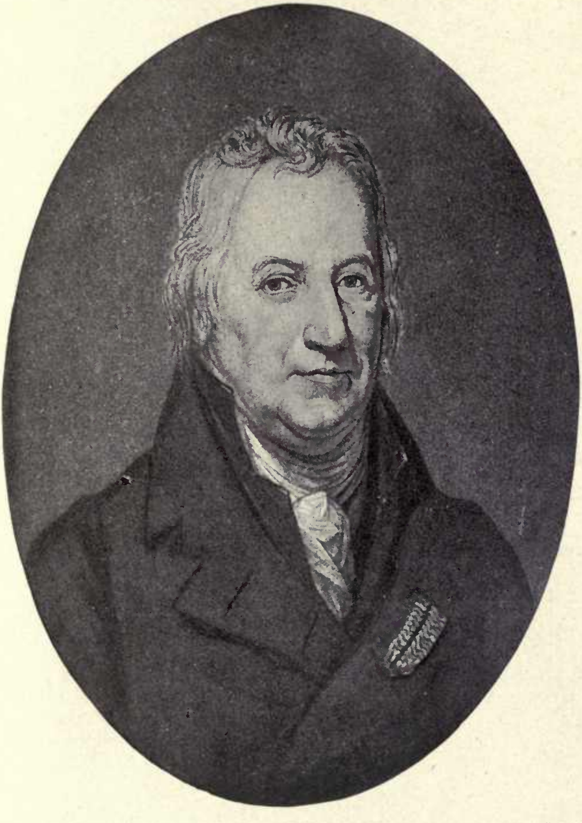
Its revelations of the gradual simplifying of elements present, as hotter and hotter stars are examined—there being only one-half of our list in the sun, and still fewer in Sirius—was taken as an encouragement to the belief that elements here unchangeable might under other circumstances be shown to be compounds.

We cannot discuss every important step made in more recent years, as again the great mass of material forbids. But since the discovery of the Röntgen rays—the x-rays, in 1905, there has been a most startling advance in the study of matter. Following Röntgen came the French chemist, Becquerel, with the discovery of " rays " from uranium, and the German, Schmidt, who found rays from thorium. Then, at Becquerel's instance, M. and Mme. Curie made a thorough examination of the elements, and found no " rays " except in these two —uranium and thorium. But rays were abundant in certain minerals—especially pitchblende, wherefrom came rays even stronger than those from uranium.

Patient, heroic work began. A car-load of pitchblende was laboriously analyzed, and out of it was secured one-quarter of a grain of chloride of "radium"—a new element. This was the $^{1}/_{4000000}$th part of the pitchblende, and the quarter grain had been evenly diffused throughout! Its cost was at the rate of $25,000,000 a kilogram.

Then began the study of the new element, and of its *three* kinds of rays—known as α, β, and γ—alpha, beta, and gamma rays. This study was exhaustive, accurate, and complete. It has been verified by observers throughout the world, and leads to most amazing conclusions. When certain of these rays or emanations are collected in a glass tube, and subjected to examination, it has been found that there is no trace of radium, but that in the tube is a gas—the gas of the element helium, as the spectrum of it proves.

Again we must condense. Let us say in brief that here are some of the chemists' conclusions: In pitchblende, where uranium is found, is found also a *proportionate amount* of radium. From the uranium, it is believed, may come the radium. Then the "rays" from radium change into helium. From helium there is a further change to polonium, another new element. Thus the chain already formed seems to show an element (uranium) giving rise to three others,—and it is thought also



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CLAUDE LOUIS BERTHOLLET
B. Savoy, 1748. D. Paris, 1822.

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that when the uranium shall have ceased to send forth its emanations, that which is left of it will be another element—this time a familiar one—the heavy, inert, metallic element, *lead.*

Such is the most recent evidence that there is hope that the elements may be simplified in number.

It is impossible not to recall the gradual steps by which, beginning with the masses of matter themselves, men have come to believe these made up of molecules,—have then been forced to seek further for the atoms that compose molecules; and are now finding in the atoms—so much else.

We have been forced to see electrons in the radio-active bodies, ions and kations in solutions, and to-day speculation is busy in trying to get within the orbits of electrons, hoping to know whether there is a central nucleus of matter—or nothing whatever!

CHEMICAL ELEMENTS

With Atomic Weights as Determined by the International Committee, 1906.

Name	Symbol	Metallic or Non-Metallic	Atomic Weight		Melting Point* C°	Discoverer and Date
			0 = 16	H = 1		
Aluminium	Al	M	27.1	26.9	654.5	Wohler, 1828
Antimony (Stibium)	Sb	M	120.2	119.3	629.5	Valentine, 1450
Argon	A	NM	39.9	39.6	——	Rayleigh and Ramsay, 1894
Arsenic	As	M & NM	75.0	74.4	500.	Schroder, 1694
Barium	Ba	M	137.4	136.4	475.	Davy and Berzelius, 1808
Beryllium (or Glucinum)	Be	M	9.1	9.03	1000*	Wohler, 1827
Bismuth	Bi	M	208.5	206.9	268.3	Valentine, 1450
Boron	B	NM	11.9	10.9	——	Davy and Gay-Lussac, 1808
Bromine	Br	NM	79.96	79.36	-7.27	Balard, 1826
Cadmium	Cd	M	112.4	111.6	320.	Herman and Stromeyer, 1817
Caesium	Cs	M	132.9	131.9	26.5	Bunsen and Kirchloff, 1860
Calcium	Ca	M	40.1	39.8	800.	Davy and Berzelius, 1808
Carbon	C	NM	12.0	11.91	——	——
Cerium	Ce	M	140.25	139.2	——	Berzelius and Hisinger, 1803
Chlorine	Cl	NM	35.45	35.18	-102.	Davy, 1810
Chromium	Cr	M & NM	52.1	51.7	——	Vauquelin, 1797
Cobalt	Co	M	59.0	58.56	1800.	Brandt, 1733
Copper (Cuprum)	Cu	M	63.6	63.1	1080.5	——
Erbium	Er	M	166.0	164.8	——	Mosander, 1843
Fluorine	F	NM	19.0	18.9	-223.	Ampere, 1810
Gadolinium	Gd	M	156.0	155.0	——	Marignac, 1886
Gallium	Ga	M	70.0	69.5	30.15	Boisbaudran, 1875
Germanium	Ge	M & NM	72.5	71.9	900.	Winkler, 1886
Gold (Aurum)	Au	M	197.2	195.7	1061.7	——
Helium	He	NM	4.0	4.0	——	Lockyer, 1868 (in sun); Ramsay, 1895 (on earth)
Hydrogen	H	NM	1.008	1.0	258.	Cavendish, 1766
Indium	In	M	114.0	113.1	176.	Reich and Richter, 1863
Iodine	I	NM	126.85	125.9	116.	Courtois, 1811
Iridium	Ir	M	193.0	191.5	1950-2500	Tennant, 1804
Iron (Ferrum)	Fe	M	55.9	55.5	1050-1587	——
Krypton	Kr	NM	81.8	81.2	-162.	Ramsay and Travers, 1897
Lanthanum	La	M	138.9	137.9	——	Mosander, 1841
Lead (Plumbum)	Pb	M	206.9	205.35	326.	——
Lithium	Li	M	7.03	6.98	180.	Arfordson, 1817. Bunsen, 1855
Magnesium	Mg	M	24.36	24.18	632.7	Davy, 1808
Manganese	Mn	M	55.0	54.6	1900*	Gahn, 1774

Mercury (Hydrargyrum)	Hg	M	200.0	198.5	-38.8	—
Molybdenum	Mo	M & NM	96.0	95.3	1900.	Hjelm, 1782
Neodymium	Nd	M	143.6	142.5	—	Welsbach, 1885
Neon	Ne	NM	20.0	19 9	—	Ramsay and Travers, 1898
Nickel	Ni	M	58.7	58.3	1600.	Crondstadt. 1751
Niobium (or Columbium)	Nb	M	94.0	93.3	1950.	Rose, 1846
Nitrogen	N	NM	14.04	13.93	-214.	Rutherford, 1772
Osmium	Os	M	191.0	189.6	2500*	Tennant, 1803
Oxygen	O	NM	16.0	15.88	-225.	Priestley, 1774
Palladium	Pd	M	106.5	105.7	1600.	Wollaston, 1803
Phosphorus	P	NM	31.6	30.77	44.25	Brandt, 1738
Platinum	Pt	M	194.8	193.3	1710*	—
Potassium (Kalium)	K	M	39.15	38.86	62.5	Davy, 1807
Praseodymium	Pr	M	140.5	139.4	—	Welsbach, 1885
Radium	Ra	M	225.0	223.3	—	Curie, 1898
Rhodium	Rh	M	103.0	102.2	1860*	Wollaston, 1804
Rubidium	Rb	M	85.4	84.8	38.5	Bunsen, 1868
Ruthenium	Ru	M	101.7	100.9	2000*	Claus, 1844
Samarium	Sm	M	150.0	148.9	—	Boisbaudran, 1879
Scandium	Sc	M	44.1	43.8	—	Nilson, 1879
Selenium	Se	NM	79.2	78.6	217.	Berzelius, 1817
Silicon	Si	NM	28.4	28.2	—	Berzelius, 1823
Silver (Argentum)	Ag	M	107.93	107.12	960.5	—
Sodium (Natrium)	Na	M	23.05	22.88	95.6	Davy, 1807
Strontium	Sr	M	87.6	86.94	600*	Davy, 1808
Sulphur	S	NM	32.06	31.83	115.1	—
Tantalum	Ta	M	183.0	181.6	—	Hatchett, 1801
Tellurium	Te	M & NM	127.6	126.6	450.	Reichenstein, 1782
Terbium	Tb	M	160.0	158.8	—	Mosander, 1843
Thallium	Tl	M	204.1	202.6	290.	Crookes, 1862
Thorium	Th	M	232.5	230.8	—	Berzelius, 1828
Thulium	Tm	M	171.0	169.7	—	Cleve, 1879
Tin (Stannum)	Sn	M	119.0	118.1	230.	—
Titanium	Ti	M	48.1	47.7	—	Klaproth, 1795
Tungsten (Wolfram)	W	M	184.0	182.6	2000*	d'Elihujar, 1781
Uranium	U	M	238.5	236.7	1500*	Klaproth, 1780
Vanadium	V	M	51.2	50.8	—	Del Rio, 1801. Sefstroem, 1830
Xenon	Xe	M & NM	128.0	127.0	-140.	Ramsay, 1898
Ytterbium	Yb	M	173.0	171.7	—	Marignac, 1878
Yttrium	Yt	M	89.0	88.3	—	Wohler, 1828
Zinc	Zn	M	65.4	64.9	420.	Paracelsus, 1520
Zirconium	Zr	—	90.6	89.9	—	Berzelius, 1824

* Many of these are not well determined.

THERMOMETER

C.	F.	
3500	6330	Temp. of electric arc.
"	"	Carbon vaporizes.
3000	5400	Temp. attained by Thermit.
2800	5072	Oxy-Hydrogen flame.
2500	4532	Osmium melts.
2225	4037	Iridium melts.
2231	4000	Heat in Bessemer Furnace.
1710	3080	Platinum melts.
1530	2731	Wrought iron melts.
1400	2552	White heat.
1371	2500	Steel melts.
1200	2192	Orange-red heat.
1100	2012	Copper melts.
1063	1981	Pure gold melts.
1050	1922	Cast iron (lowest) melts.
970	1778	Silver melts.
700	1292	Dull red heat.
625	1157	Aluminium melts.
405	762	(about) Coal ignites.
400	752	Red-hot iron visible in dark.
357	674	Mercury boils.
316	600	Lead melts.
288	550	Gunpowder ignites.
215	420	Tin melts.
109	228	Sulphur melts.
100	212	Water boils.
79	174	Alcohol boils.
65	149	Fusible Alloy melts.
61	142	Beeswax melts.
46	114	Paraffin melts.
44	111	Phosphorus melts.
40	104	
37.7	100	
36.8	98	Human body in health.
35	95½	
30	86	
25	77	
20	68	
17	62	Mean temp. of sea.
15	59	
10	50	Mean temp. of air (London).
5	41	
0	32	Water freezes.
-10	14	
-17.7	0	Mixture salt and ice.
-20	-4	
-38.8	34.3	Mercury freezes.
-40	-40	
-55	-68	*Greatest natural cold on earth (antarctic).
-70	-94	Sounding balloon (9 miles high).
-100	-148	
-191	-312	Air liquefies under normal pressure.
-252	-422	Hydrogen liquefies.
-257	-432	Hydrogen freezes.
-260	-436	Greatest artificial cold produced (Dewar).
-278	-459	Absolute zero.

ELEMENTS IN CRUST OF EARTH

WATER

By combining a large number of analyses of rocks of all sorts, F. W. Clarke has estimated the relative amounts of elements in the crust of the earth:—

Oxygen	47.02	per cent.	Manganese	.07	per cent.
Silicon	28.06	"	Sulphur	.07	"
Aluminium	8.16	"	Barium	.05	"
Iron	4.64	"	Strontium	.02	"
Calcium	3.50	"	Chromium	.01	"
Magnesium	2.65	"	Nickel	.01	"
Sodium	2.63	"	Lithium	.01	"
Potassium	2.32	"	NM. Chlorine	.01	"
Titanium	.41	"	NM. Fluorine	.01	"
Hydrogen	.17	"			
Carbon	.12	"		100	
Phosphorus	.09	"			

—*Science Year Book*, 1908.

Mean density of the whole earth is 5.53 times that of water.
—*Science Year Book*, 1908.

* Capt. Amundsen reported (in 1905) a temperature of —61.7° C (or — 79° F) in Boothia (N. Canada). — 68° C is said to have been experienced in Siberia.

Fahrenheit derived his scale from putting zero as the greatest cold then ascertained (by mixing salt and ice). The temperature of the human body was the other standard, and the space between these points was divided first duodecimally into 24° and later into one quarter of this, or 96°.

Reaumur's scale (used in Germany) divides the space between the freezing and boiling points of water into 80°.

To convert these scales—

$$F^\circ = \frac{9}{5} C^\circ + 32 = \frac{9}{4} R^\circ + 32$$

$$C^\circ = \frac{5\ (F-32)}{9}, \qquad R^\circ = \frac{(4\ F-32)}{9}$$

THE AIR

1 cubic foot of air at 62° F. weighs .076 lbs. (= 1.217 ozs. or 532.7 grains); at 32° F., .08 lbs. 1 litre at 32° F. weighs 1293 grammes. 13.141 cubic feet at 62° F. weigh 1 lb.

Normal atmospheric pressure 14.7 lbs. per sq. inch = 2116.4 lbs. per sq. foot. A column of mercury 30 inches high or a column of water 33.947 ft. is supported.

Sir F. Abel and Sir A. Noble obtained (in researches with explosives) pressures in closed steel cylinders up to 95 tons to the square inch.

Air expands, or contracts, .002 of its volume per ° F.

Air, under normal pressure, liquefies at—191° C. At 39 atmosphere pressures it will liquefy at—140° C.

Liquid air occupies $\frac{1}{800}$th the volume of air.

The average composition of normal air may be taken as follows:—

	Vols.
Nitrogen	769.5000
Oxygen	206.5940
Aqueous vapour	14.0000
Argon	9.3700
Carbon dioxide	0.3360
Hydrogen	0.1900
Ammonia	0.0080
Ozone	0.0015
Nitric acid	0.0005
	1,000.0000

Also—

Neon	0.000014
Helium	0.00000276
Krypton	a trace
Xenon	a trace

SIZE OF MOLECULES

The diameters of Molecules have been ascertained by Jeans to be—

Hydrogen	20.3
Nitrogen	29.1
Oxygen	27.3

These figures express number of billionths of a metre.

METALLIC SALTS

These are formed when metals replace the hydrogen of acids; or when acid-forming oxides combine with basic oxides; or when metals exchange with hydrogen in combining an acid and a hydroxide.

If all the hydrogen is replaced, "normal salts" are formed; if part is left, an "acid salt."

Basic salts result from the combining of a normal salt with a basic oxide or hydroxide.

CONDUCTIVITY OF METALS

Substance.	Heat Cond.	Electrical Cond.
Silver	100.0	100.0
Copper	73.6	73.3
Gold	63.2	58.5
Aluminium	31.3	60.5
Brass	23.6	21.5
Zinc	19.9	29.0
Tin	14.5	22.6
Iron	11.9	13.0
Lead	8.5	10.7
Platinum	6.4	10.3
Bismuth	1.8	1.9
Mercury	—	1.6

WATER

Composition:—2 vols. Hydrogen to 1 Oxygen. Maximum density at 4° C. (below which it *expands* slightly).

Water converted to steam expands to 1700 times its volume.

Water is 819.4 times as heavy as air.

Composition of rain water (London):—

Organic Carbon...	.99	part	in	1,000,000	water.
Organic Nitrogen.	.22	"	"	"	"
Ammonia	.50	"	"	"	"
Nitrates and Nitrites	.07	"	"	"	"
Chlorine	6.30	"	"	"	"
Total Solids......	39.50	"	"	"	"

COMMON NAMES OF CHEMICALS

Common Names.	*Chemical Names and Formulae.*
Alum	Sulphate of Aluminium and Potassium
Aqua Fortis.................	Nitric Acid, HNO_3
Aqua Regia.................	Nitro-Hydrochloric Acid
Calomel	Mercurous Chloride, $Hg_2 Cl_2$
Carbolic Acid................	Phenol C_6H_5OH
Caustic Potash...............	Potassium Hydrate KOH
Caustic Soda................	Sodium Hydrate, NaOH
Chalk	Calcium Carbonate, $CaCO_3$
Copperas	Sulphate of Iron
Corrosive Sublimate..........	Mercuric Chloride, $HgCl_2$
Cream of Tartar.............	Potassium Bitartrate
Epsom Salts.................	Magnesium Sulphate
Ether	Diethyl Oxide $(C_2H_5)_2O$
Fire Damp..................	Light Carburetted Hydrogen
Galena	Lead Sulphide, PbS
Glauber's Salt...............	Sodium Sulphate
Glucose or Grape Sugar......	Dextrose $C_6H_{12}O_6$
Goulard Water..............	Basic Acetate of Lead
Iron Pyrites.................	Iron Di-Sulphide, FeS_2
Jewellers' Putty.............	Oxide of Tin
Laughing Gas................	Nitrous Oxide, NO

COMMON NAMES OF CHEMICALS—*Continued*

Lime	Calcium Oxide, CaO
Lunar Caustic	Silver Nitrate, $AgNO_3$
Mosaic Gold	Bi-Sulphide of Tin
Muriatic Acid	Hydrochloric Acid, HCl
Olefiant Gas	Ethylene, C_2H_4
Plaster of Paris	Calcium Sulphate
Quartz	Silicon Dioxide, $Si\ O_2$
Realgar	Arsenic Di-Sulphide, $As_2\ S_2$
Red Lead	Oxide of Lead, $Pb_3\ O_4$
Rochelle Salt	Sodium Potassium Tartrate
Salammoniac	Ammonium Chloride
Salt, Common	Sodium Chloride $NaCl$
Salt of Tartar	Potassium Carbonate
Saltpetre	Potassium Nitrate, KNO_3
Salts of Lemon	Oxalic Acid
Slaked Lime	Calcium Hydrate
Soda	Sodium Carbonate
Spelter	Zinc
Spirits of Hartshorn	Amm. Hydroxide, $N_4H.OH$
Spirits of Salt	Hydrochloric Acid, HCl
Sugar of Lead	Lead Acetate
Tartar Emetic	Potass. Antimony Tartrate
Verdigris	Basic Copper Acetate
Vermilion	Sulphide of Mercury
Vinegar	Dilute Acetic Acid
Vitriol, Blue	Copper Sulphate
“ Green	Ferrous Sulphate
“ Oil of	Sulphuric Acid, H_2SO_4
“ White	Zinc Sulphate
Volatile Alkali	Ammonia

FREEZING MIXTURES

	Ingredients.	Parts by Weight.	Diminution of Temperature.
1	Snow or pounded ice......... Chloride of sodium (salt)....	5 2	20°C.
2	Water .... Saltpetre .... Chloride of ammonium (sal ammoniac)	16 5 5	22°
3	Water .... Nitrate of ammonia..........	1 1	26°
4	Water .... Nitrate of ammonia.......... Carbonate of soda............	1 1 1	29°
5	Snow .... Crystallized chloride of calcium	1 2	45°
6	Crystallized sulphate of soda. Hydrochloric acid............	8 5	50°
7	Solid Carbonic Acid dissolved in Sulphuric ether..............	—	146°

EXPLOSIVES

GUNPOWDER (English)—

Saltpetre	75	parts.
Charcoal	15	"
Sulphur	10	"

Ground to fine powder and intimately mixed: then granulated. 1 c. in. on exploding expands to 800 c. in. of gas.

Guncotton—Cotton immersed in three parts sulphuric acid (sp. gr. 1.84) and 1 part nitric acid (sp. gr. 1:52) and thoroughly washed; 2½ to 4 times as powerful as gunpowder.

Nitro-Glycerine ($C_6 H_5 N_3 O_{18}$), sp. gr. 1.6—

Nitric acid......	1½ parts by wt.
Sulphuric acid...	2 " "

Glycerine being injected into the mixture, nitro-glycerine floats to the top, is drawn off, washed, and filtered.

Explodes at 240° to 300° F., 1 c. in. expanding to 10,000 c. in. of gases, or about 10 to 13 times as strong as gunpowder.

Dynamite—3 parts by wt. nitro-glycerine, mixed with 1 part Kieselguhr. (infusorial earth).

Blasting Gelatine—Guncotton dissolved in nitro-glycerine.

Cordite—Practically the same.

Picric Acid (Lyddite, Melinite, etc.)—Carbolic acid treated with nitric acid.

CHEMICAL NOMENCLATURE.

Termination "—UM" is now applied to all **Metals**, though the older-known metals retain the former names, *e. g.*—Aluminium, Tellurium, etc.

Termination "—IDE" denotes a **Binary Compound**, that is, a substance composed of only two elements, *e. g.*—Sodium Chloride (NaCl).

Termination "—OUS" is applied to the first of two elements when it exists in a greater proportion than in another combination with the same element, *e. g.*—one atom of phosphorus and three atoms of chlorine form **Phosphorous Chloride.**

Termination "—IC," when the first exists in a lesser proportion, *e. g.*—one atom of phosphorous with five atoms of chlorine form **Phosphoric Chloride.**

Prefixes —"MONO—," "BI—," "TRI—," &c., indicate the proportion of the latter of two elements, and are sometimes used instead of the above termination. Thus phosphorus chloride may also be called **Phosphorus Tri-Chloride**; so one atom of carbon with one atom of oxygen is **Carbon Monoxide.**

Prefix —"HYPO—" (under) and "PER—" (over), specify compounds formed by the same two elements containing less (or more) of an element than is in the usual compound.

Sesqui—means in the proportion of 1 to 1½ or 2 to 3.

ALLOYS

Bell Metal = 4 copper, 1 tin.
Brass = 1 zinc, 2 copper.
Bronze (coins) = 95 copper, 4 tin, 1 zinc.
German Silver = 3 copper, 1 nickel, 2 zinc.
Gold (coinage = 22 carats)* = 11 gold, 1 copper.
Gold (Jewellers' = 18 carats) = 3 gold, 1 copper.
Gun Metal = 9 copper, 1 tin.
Invar = 2 steel, 1 nickel.
Magnetic Alloy = 60 copper, 26 manganese, 14 aluminium.
Pewter = 4 tin, 1 lead.
Silver (coinage) = 37 silver, 3 copper.
Solder (soft) = lead and tin (varies).
Solder (hard) = copper and zinc (varies).
Speculum Metal = 126.4 copper, 58.4 tin.
Type Metal = 4 lead, 1 antimony, and tin (var.)

WOOD'S FUSIBLE ALLOY

Bismuth, 4 parts.
Lead, 2 parts
Tin, 1 part.
Cadmium, 1 part.

Melts at 149° F.

METALLIC OXIDES

Monoxide = replacing of each atom of hydrogen in H_2O by a monad, or both by a dyad.

Higher oxides = replacing of hydrogen atoms in molecules of water by equivalent atoms of metals.

Hydroxides = only a part of hydrogen is replaced in the water molecules. In solutions with water these have alkaline reaction.

Basic oxides and hydroxides form salts with acids, the metal replacing hydrogen.

Peroxides, and Acid-forming Oxides, have more oxygen than basic oxides.

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* Pure gold has 24 carats (of 240 grains) in a lb.

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